

A JUNIOR CHEMISTRY

W. D. ROGERS





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BY W. D. ROGERS

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JUNIOR CHEMISTRY

BY W. D. ROGERS

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JOHN MURRAY, ALBEMARLE STREET, W.

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BY W. H. BAKER

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1910

PREFACE

THIS book is based upon an experience of teaching young pupils which has extended over many years. It is an attempt to present a thoroughly sound course of experimental chemistry on modern lines, and though it was originally designed to fit the syllabus of the Junior Examination of the Central Welsh Board, it has been expanded, at the publisher's request, so as to cover the ground needed for all kindred examinations of a similar standard.

The course is divided into four parts, each of which forms a complete and separate unit. It is based throughout upon laboratory work, and while the text itself provides full material for class instruction, the author has added a large number of exercises at the end of each part, from which extra work may be taken at the discretion of the teacher. This material consists of (1) a set of questions for discussion; (2) a series of test-papers dealing with both theoretical and practical work, consisting largely of questions actually set in the various public examinations.

Short historical sections are included in smaller type, with the object of stimulating interest, and an attempt has been made to render free discussion and revision a habit. Queries are inserted broadcast, and cross-references to earlier work are plentiful. The apparatus described is of the simplest possible type consistent with accuracy. This seems better adapted to young pupils, and it possesses the additional merit of being more easily reproduced from memory. All the diagrams (with the exception of Fig. 55¹) have been specially prepared for this book, and simplicity of design has been the first consideration.

For pupils who take only one branch of science, and

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that chemistry, it has been deemed advisable to include an elementary chapter on Heat; and some general instruction in elementary Physics is also distributed throughout the book.

Equations of changes have not been introduced until they are fully understood, and the letters in square brackets in the text [A] refer to a special chapter on Equations (Chapter XXVI.). This is a novel feature in textbooks, and it is hoped that it will be made the means of systematically revising earlier work, and not merely the occasion of memorizing a list of formulæ.

A word of explanation is perhaps needed about the position in the book of the Balance. It is put at the beginning, in deference to prevailing opinion; but it is separated from the text. The author's view is that its use by quite young pupils is to be deprecated: it is an expensive piece of apparatus, and, what is more, too early efforts at weighing are apt to be incorrect, and to lead to wrong conclusions. The historical development of the subject has therefore been followed: at first, the instructor alone uses the balance, in the first two chapters the pupils' work is qualitative only, but afterwards a quantitative course is followed throughout.

The author would like to acknowledge his indebtedness to those whose encouragement has resulted in the publication of this little book, especially to Mr. W. Hammond Robinson, whose kindly interest and helpful criticism have been invaluable during its entire preparation.

To the various examining bodies which have kindly given permission to reproduce questions, he would also like to express his thanks: to the London University Matriculation Board, the Delegacies of the Oxford and Cambridge Local Examinations, the College of Preceptors, and the Central Welsh Board. The individual questions taken from their papers are identified by initials in each case.

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CHEMISTRY

THE BALANCE.

WHEN an experiment is carried out in order to find merely the kind of action occurring, or the nature of the bodies reacting, and so on, it is termed **qualitative**. Such was the work done in chemistry before the eighteenth century, and often led to many absurd beliefs. In 1775 the famous French chemist, **A. Lavoisier** (1737-1792), pointed out the importance of accurate weighings being taken at every stage of a reaction, and since his time the chemical balance has been the guide of the chemist in his work. Such work is **quantitative** (or **gravimetric**).

To examine a Chemical Balance.—Examine your balance, and compare it with the diagram (Fig. 1). It consists of a beam (*B*), fitted with three knife-edges (*e*), set at right angles to the beam length. These are made of hard steel or agate. The central knife-edge rests on a plane or in a notch of the same material (*p*) on a rod (*r*), which is moved up and down inside the pillar (*P*) by turning the lever handle (*L*). The pillar carries two fixed side-arms (*AA*), on which the beam rests when the balance is not in use, the knife-edges being then off their planes. (Why ?) At the lower

end of the pillar a graduated scale (*S*) is fixed to render easily visible the movement of the pointer (*i*), which is fastened, at right angles to the beam, in front of the central knife-edge. The scale-pans (*g*) are suspended in stirrups (*h*), which carry planes, and rest on the end knife-edges of the beam. Usually two small mill-heads (*m*), travelling on screws, are attached at the

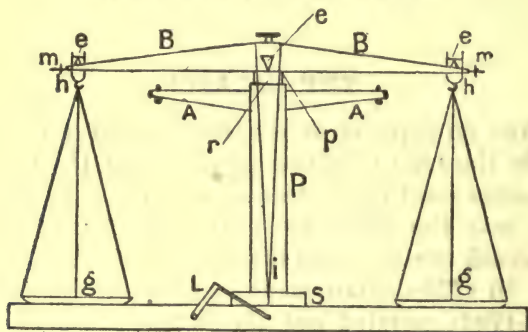


FIG. 1.

ends of the beam for adjusting the balance, and the instrument is enclosed in a glass case to avoid draughts, etc.

To examine a Box of Weights.—You will observe that your weights are arranged as in the following manner: (1) A set of brass or gilded weights; (2) a set of smaller weights preserved under a glass plate. (Why?) In addition, a pair of forceps is provided to avoid using the moist fingers in weighing. Note that the weights occur in the following order :

50 gms.	20 gms.	20 gms.	10 gms.
5 gms.	2 gms.	2 gms.	1 gm.

The smaller weights are in the same order; thus:

500 mgms.	200 mgms.	200 mgms.	100 mgms.
or	or	or	or
0.5 gm.	0.2 gm.	0.2 gm.	0.1 gm.
50 mgms.	20 mgms.	20 mgms.	10 mgms.
or	or	or	or
0.05 gm.	0.02 gm.	0.02 gm.	0.01 gm.

Precautions to be observed when using a Balance.—

The balance is a delicate instrument, and must not be treated carelessly.

1. When the balance is at rest, the pointer should be at the zero of the scale.

2. The balance when not in use should always be left in the *position of rest*, to avoid the useless wearing of the knife-edges. (What effect has this ?)

3. Nothing must be placed on or taken off the pans while the balance is swinging. (Why ?)

4. Place the weights on the right-hand pan with the forceps. Never touch the weights or object to be weighed with the fingers. (Why ?)

5. When weighing, first find a weight which is too heavy; remove it and try the next smaller weight, and so on, in regular descending order. Never try to *guess* the required weight.

6. Before removing the weights, add them up and check your result as you replace them in the box. Your rough notebook should appear thus:

Weight of empty dish:

10	or	10
2		2
1		1
0.5		0.500
0.02		0.020
0.01		0.010
<hr/>		<hr/>
13.53 gms.		13.530 gms.

7. Never weigh anything that is hotter than the air, or your weighing will be inaccurate, and you will injure the balance-pan.

8. Before weighing, test your balance by observing whether it oscillates equally on both sides of the scale zero. If it does not, it must be adjusted by moving the mill-heads. The balance must swing over an equal number of scale divisions on each side, and not "stick" at the centre, for an accurate weighing.

EXERCISES.—(1) Weigh a penny, an evaporating-dish, and a watch-glass. (2) Find whether any change in weight occurs during the physical and chemical changes. Exercise 1, (1), (2), (5), (6), Chapter II., using small quantities in heating-tubes, and note what conclusions you arrive at.

THE METRIC SYSTEM OF WEIGHTS AND MEASURES.

This system has one main advantage—viz., its convenience. It is a decimal system, the only multiple or submultiple employed being 10, instead of, as in the English system, 16, 14, 20, $5\frac{1}{2}$, $30\frac{1}{4}$, etc.

The units in this system are—

- (i.) of weight—the **kilogramme (kgm.)**
= 1000 grammes = $2\frac{1}{5}$ lbs. (av.) approx.
- (ii.) of length—the **metre (m.)**
= 39.37 inches = 1.093 yards approx.
- (iii.) of volume—the **litre (l.)**
= 1000 cubic centimetres = 1.75 pints approx.

Special prefixes are used to denote multiples and sub-multiples of these units; thus:

Deca- (D.) = 10. *e.g.*, a decametre = 10 metres.
 Hecto- (H.) = 100. *e.g.*, a hectolitre = 100 litres.
 Kilo- (K.) = 1000. *e.g.*, a kilometre = 1000 metres.

deci- (d.) = $\frac{1}{10}$, or 0.1. *e.g.*, a decigramme
 = 0.1 gramme.

centi- (c.) = $\frac{1}{100}$, or 0.01. *e.g.*, a centimetre
 = 0.01 metre.

milli- (m.) = $\frac{1}{1000}$, or 0.001. *e.g.*, a millilitre
 = 0.001 litre.

Hence, 10 gms. (m. or l.) = 1 D.-gm. (m. or l.).

10 D.-gms. (m. or l.) = 1 H.-gm. (m. or l.).

10 H.-gms. (m. or l.) = 1 K.-gm. (m. or l.).

and 10 m.-gm. (m. or l.) = 1 c.-gm. (m. or l.).

10 c.-gm. (m. or l.) = 1 d.-gm. (m. or l.).

10 d.-gm. (m. or l.) = 1 gm. (m. or l.).

The above units are too large for most scientific purposes, and for these the following smaller units are chosen:

- (i.) of weight—the **gram (gm.)** = $\frac{1}{1000}$ Kgm.
- (ii.) of length—the **centimetre (cm.)** = $\frac{1}{100}$ metre.
- (iii.) of volume—the **cubic centimetre (c.cm.)**
 = $\frac{1}{1000}$ litre.

The **gram** may be “roughly” defined as the weight of 1 c.cm. of water at 4° C. It is useful to remember the following approximate English equivalents:

1 gm. = 15.432 grains. 1 oz. (av.) = 28.4 gms.

1 fl. oz. = 28.4 c.cm.

30.4 cms. = 1 foot. 2.54 cms. = 1 inch.

Moreover, the prefixes above mentioned are rarely used—*e.g.*, we prefer to write 532.47 gms. rather than 5 Hgms. 3 Dgms. 2 gms. 4 dgms. 7 cgms. The only ones of any importance being **kilo-** (= 1000); **centi-** (= $\frac{1}{100}$); and **milli-** (= $\frac{1}{1000}$).

QUESTIONS.

1. Of what use do you think the chemical balance is to the chemist? Give examples of changes in which the use of the balance has pointed out errors and shown the correct conclusion.
2. Write a short essay on the construction of the balance, and point out the use of the various parts.
3. Describe a balance such as you have actually seen, and name the various parts of it. Illustrate your answer by sketches.
4. Describe carefully all the precautions you would take in order to weigh an object accurately.
5. If 1 inch = 2.54 cms., find (1) the number of cms. in 1 mile; (2) the number of Kms. in 1 mile.
6. If 1 inch = 2.54 cms., find (1) the number of c.cms. in 1 c. foot; (2) the number of litres in 1 c. foot.
7. If 1 gram = 15.432 grains, find (1) the number of grams in 1 lb. (av.); (2) find the number of Kgms. in 1 lb. (av.).
8. If $28\frac{1}{2}$ grams = 1 ounce, find the number of mgms. in 1 lb. (av.).
9. Express as decimals of a gram (v.v.): (1) 500 mgms., 530 mgms., 270 mgms.; (2) 30 mgms., 75 mgms., 195 mgms.; (3) 55 cgms., 35 cgms., 85 cgms., (4) 45 dgms., 27 dgms., 75 dgms.
10. Express the following in mgms., cgms., and in dgms. (v.v.): 0.523 gm.; 0.475 gm.; 0.625 gm.; 0.074 gm.; 0.005 gm.; 0.108 gm.
11. Express in grams the following (v.v.): (1) 1.5 Kgms., 37.75 Kgms., 0.65 Kgm.; (2) 2.7 Hgms., 32.5 Hgms., 172.6 Hgms.; (3) 57.5 Dgms., 6.75 Dgms., 532.7 Dgms.
12. Express the above weights (Question 11) in Kgms., Hgms., Dgms. (v.v.).

INSTRUCTIONS FOR PRACTICAL WORK.

1. Keep your bench clean and tidy. A dirty bench means slovenly, inaccurate work. Should any liquid or solid be spilled, it should be wiped up **at once**. Any apparatus, etc., finished with should be cleaned and put away immediately.

2. Before commencing an experiment, read the account of it, and be quite certain that you understand the object of it, how you are to carry it out, and what apparatus you require.

3. Observations must be noted down in your notebook **at once**, and hence in the order in which they occur; do not trust your memory. Do not write down what you think ought to occur, but what actually takes place. All weighings must be noted down immediately at the balance before removing the weights from the scale-pan.

4. Every experiment **must teach you something**—*i.e.*, must have a **conclusion**.

5. When washing a substance on a filter, etc., make sure that the washing is complete—*i.e.*, test the last spot. **Remember** frequent washing with **small** quantities of **hot** liquid is much better than fewer washings with larger amounts of cold liquid.

6. When weighing anything increasing or decreasing in weight, make sure that the operations are complete—*i.e.*, repeat until two consecutive weighings agree or **until of constant weight**.

7. Remember that the hot part of the Bunsen flame is the upper portion. Look and see that the portion to be heated is hot, and not the portion above or below it.

8. Glass vessels containing liquids to be heated must be dry on the outside and heated on the wire-gauze or sand-tray. The flame must be adjusted so that it does not play above the surface of the liquid.

9. In fitting up apparatus, remember that all corks must be first **softened**. They must be bored from the narrow end. All glass tubes must be rounded off, and must be drawn out or blown **out of the flame**. Cold glass must never be **suddenly** heated nor hot glass **suddenly** cooled.

10. Be quite certain that all apparatus for collecting gases is air-tight before commencing the experiment. Close all the holes save one, and draw out a little air. If air-tight, suction on the tip of the tongue will occur.

11. When collecting gases and measuring their volume, arrange for cooling to occur. See that you measure **all** the gas, neither more nor less.

12. Every weighing or measurement obtained must have the name of the substance weighed or measured. This avoids confusion and often the consequent failure of your experiment (see Appendix).

PART I

CHAPTER I

MATTER—ENERGY—HEAT

What is Matter?—The various bodies around us—earth, stones, metals, water, etc., even the air—all have certain properties or characteristics in common—*e.g.*, they all possess weight. The word **matter** (stuff or substance) will be used in this book to mean **anything that has weight**.

The States of Matter.—Common experience tells us that matter exists in three physical conditions, for we know of **solids, liquids, and gases**. The two latter kinds of matter are sometimes collectively known as **fluids**, thereby denoting that they *flow* from place to place, as distinct from solids that remain in the same position if left to themselves. We do, however, know of solids (?) which flow extremely slowly—*e.g.*, a piece of sealing-wax or pitch supported at either end “sags” in the middle, and obviously resemble viscous liquids, like tar or treacle.

Solids, then, are distinct from other types of matter in the following points:

1. They possess a **definite shape and volume**.
2. They resist and recover from any change of **shape or volume**—*i.e.*, they possess **elasticity and rigidity**.

3. They have **great cohesion**—*i.e.*, to separate the particles of a solid requires the exercise of great force.

4. They **transmit a force in one direction only**—*e.g.*, if we push against a wall with a stick, we do not find any force exerted at the side of the stick.

5. Their **densities**—*i.e.*, the weight of a unit volume—are **greater** than the density of matter in other states.

Fluids differ from solids in that they—

1. Have **no definite shape**, but merely take the shape of the containing vessel.

2. Have **only elasticity of volume**.

3. Have **only small cohesion**—*e.g.*, they readily flow; *i.e.*, their particles easily move.

4. **Transmit pressure equally in all directions**.

Liquids are distinguished from gases by their having an almost invariable volume—*i.e.*, a **high volume elasticity**. For a long time liquids were thought to be incompressible. Consequently they *only need support at the base and sides*, and *fill the lower portion only* of the vessel in which they are placed.

Gases, on the other hand, are characterized by the apparent **absence of cohesion**; they **flow readily** and rapidly, **mix in all proportions**, are **easily compressed**—*i.e.*, have a *low volume elasticity*—and *fill the whole* of the vessel in which they are contained—hence they *need support on all sides*.

Is Matter always in the Same State or Condition?—A little thought will soon convince us that it is not. The rock perched on the hillside to-day is found in the valley to-morrow. The same rock—*i.e.*, the same matter. Why has it moved? Whatever it was that caused the rock to fall cannot be seen or traced in any

way, so that it cannot be matter, whatever else it may be. How is it that we can draw water so easily from a tap instead of laboriously drawing it from a well? It is obviously the same matter—water—but in a different condition. The poker to-day is cold, yesterday it was hot. Is it not the same poker in every way? Why, then, does it cause so different a sensation? A boy's arm when striking a blow—although the effect may be great—is still the same arm. What causes the effect? An electric lamp only gives out light while electricity is passing, yet it is quite unchanged in material throughout. What is this peculiar "electricity"? Is it matter or not? What is "light"? Can it be weighed like matter?

It thus seems clear that, in addition to matter, there is something else we must concern ourselves with (which is quite different from matter) that causes manifold changes in its state or condition. It causes matter to become hot or cold, to give out light, to do work, and so on. To this peculiar something we give the name **energy**, and it is made known to us in various ways, as **force, heat, light, sound, electricity, etc.** **Heat** may therefore be defined as **that form of energy which gives us the sensation of hotness and coldness.**

EXPT. 1.—To show that Heat is not Matter.—Take two equal cylinders of copper and place them in wooden boxes containing cotton-wool. Counterpoise them with sand or weights. Place one of them before a fire or in a steam oven, and when hot, replace it in its box, close its lid, and see whether it remains balanced or not. Does not your experiment prove that heat has no weight?—*i.e.*, **heat is not matter.**

Since we cannot weigh heat, we must study its effect

on various substances to discover some means of measuring it.

EXPT. 2.—To study the Effect of Heat on Solids.—

1. Heat fragments of iron, lead, tin, magnesium, platinum, sulphur, charcoal, etc., in iron spoons or crucibles, and observe that **heat causes some solids to melt, some to burn, whilst others merely become hot for a time.**

2. Arrange an iron retort-stand as shown (Fig. 2). *A* is a wood block on which is placed a needle (*n*), having attached to it a straw or wood spill (*p*), to act as pointer. The end of the retort-stand is laid gently

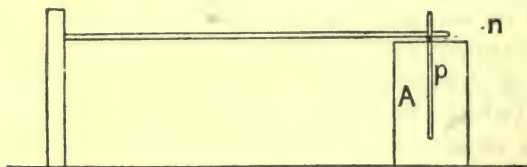


FIG. 2.

on the needle, and the pointer adjusted to the vertical. Now heat the rod strongly and observe the pointer. Allow the bar to cool, and watch as before. If this experiment is repeated with rods of different materials and of equal lengths, heating all equally, the same change is observed, but the amount of movement of the pointer varies. We may thus conclude (1) **heat causes solids to expand**; (2) **different solids expand unequally.**

These facts are very important, and have to be taken note of in metal constructions. Many of you will no doubt have noticed that railway-lines are laid in sections, each section being separated from the next, and you will now know why this is done.

Before proceeding further, it is necessary to be able to perform a few simple laboratory operations.

EXPT. 3.—To Cut Glass Tubing.—Procure a good triangular file and a piece of ordinary glass tubing. Draw the file *smartly across* the tubing *towards you*, with moderate pressure (too great a pressure will crush the tube), once *only*, making a clean scratch across the glass. Hold the glass tube between the thumb and finger on either side of the file-mark, and with the mark away from you, bend the ends *slightly* towards you, and at the same time pull them *gently* apart. The glass should snap cleanly across at the file-mark, if this is deep enough. **Be gentle**; if you use much force, the glass will splinter, making a jagged edge, and probably cut your hands. If it does not snap easily, make a deeper file-mark and repeat.

EXPT. 4. — To “round off” the Ends of Glass Tubing.—If you examine the ends of the tube cut off in Expt. 3, you will find that they are quite sharp, and would, of course, cut any cork they were pushed through, and also your hands. To avoid this, the end should *always be rounded off* in the following manner: Rotate the tube between the thumb and fingers, and bring it slowly into the tip of the Bunsen flame. Keep it there, still rotating, until the flame is markedly yellow. Remove it and examine. You will find that the jagged appearance has disappeared and a smooth surface remains. **Caution.**—*Do not keep the tube in the flame too long, or you will contract the bore, and even close it completely.*

EXPT. 5.—To bore a Cork.—You are provided with a series of metal tubes and a handle. Choose a borer with bore just *smaller* than the glass tube supplied,

and see that its cutting edge is *sharp*. Place the handle through the holes at the top and practise boring on old corks until you can bore a *straight, clean hole* with a *clean exit*. Bore pressing *gently* and with a *rotating motion*, clearing out the cork in the borer as it gets clogged. When nearing the other side of the cork, place it *against the bench* and *cut* a clean exit; do not push out a jagged lump. When you can bore satisfactorily, then try on a new cork. First choose a cork a little too *large* for your purpose, and *soften* it by squeezing in a cork-press or rolling it gently under your foot. Secondly, always *begin* boring from the *narrow end*. (Why?)



FIG. 3.

EXPT. 6.—To study the Effect of Heat on Liquids.—1. Place a small quantity of water in an evaporating-dish and heat this on a wire-gauze, supported on a tripod stand, with the Bunsen flame. Place small quantities of methylated spirits, chloroform, etc., in separate dishes on a steam oven. Observe them from time to time, and verify the statement that **heat causes liquids to boil away or evaporate in the form of vapour or gas.**

2. Fill several small flasks of the same size with different liquids—coloured with a little ink—such as those mentioned above, and attach a cork provided with a long tube, as shown (Fig. 3). Fix a scale ruled on paper behind each tube and observe the level of each liquid. Now put them all on a steam oven or in a pot of hot water, and closely observe the liquid level. Do not these experiments show clearly that **liquids expand when heated**, and, moreover, they **expand more than solids?**

If you observe very closely, you will find that at first the level falls—this is due to the expansion of the *solid* glass—and then rapidly rises as the liquid becomes heated. If your flasks, etc., are all of the same size, you may also notice that **different liquids expand unequally and at different rates.**

EXPT. 7.—To observe the Effect of Heat on Gases.—Fit up the apparatus shown in Fig. 4, which consists of a flask, fitted with a cork, through which passes a tube, having attached to it a scale as used in Expt. 6. Place the hands on the flask, bring a candle near to it, etc., and observe the very **rapid, large** expansion of the gas. Repeat the experiment, filling it with coal-gas, and verify the fact that **heat causes gases to expand much more than liquids.** Careful experiment also shows that **all gases expand equally** when heated to the same extent.

Thus the general effect of heat on all kinds of matter is to make them expand. In order to make use of this effect in the investigation of heat, we must have a large enough expansion to be easily measured (solids will not be suitable therefore), and not too large to be convenient. Consequently liquids are usually employed. We shall now investigate the expansion of liquids more closely, and for the experiment it is necessary to make a glass bulb at the end of a glass tube.



FIG. 4.

EXPT. 8.—To blow a Bulb of Glass.—A piece of moderately small-bored (capillary) tube is chosen of uniform bore. Cut off, as in Expt. 3, a piece about

15 inches long, and first warm the tube thoroughly by moving it rapidly in and out of the flame. (This prevents the tube from cracking.) Now hold the end of the tube in the tip of the Bunsen flame, *rotating it slowly*, until the end is quite closed and the tube appears as at *A*, not as at *B*. Proceed now to heat the end $\frac{1}{4}$ inch, rotating it continuously to heat all sides equally, and when red-hot all through, *remove it from the flame* and blow *gently* in a series of puffs.

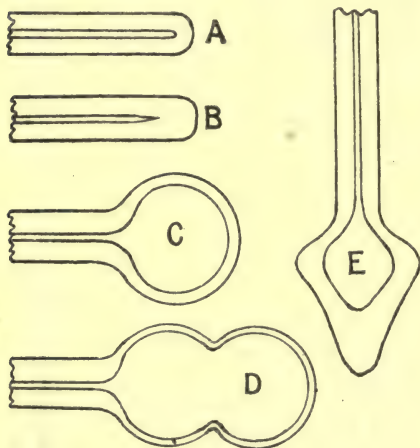


FIG. 5.

It will then appear as at *C*. Heat the next $\frac{1}{4}$ inch, and by again blowing produce *D*. Now use a large flame, and, rotating rapidly, heat the expanded portion, holding the tube almost vertically. When it has the appearance as at *E* (pear-shaped), remove it from

the flame, and whilst holding it vertically, blow *gently* and *steadily*, rotating all the time. A perfectly spherical bulb of uniform thickness and of suitable size ($\frac{1}{2}$ to $\frac{3}{4}$ inch in diameter) should be formed. If not, repeat the last heating and blowing processes. Test your bulb by dropping it from a height of 3 inches on to the bench, allowing the stem to slip loosely through the thumb and forefinger. The bulb, if a

sound one, will remain unbroken. **Caution.**—*Glass must never be bent or blown whilst still in the flame, nor be kept stationary there.*

EXPT. 9.—To study the Expansion of Liquids.—Fill the bulb with boiled distilled water in the following manner. Attach a short thistle funnel to the open end of the bulbed tube by means of india-rubber tubing, and suspend it on a retort-stand. Place in the thistle funnel some of the liquid which has been heated to a moderate extent. Why does it not run into the bulb? Gently *warm* the tube and bulb, and observe how the air escapes through the liquid. Do not heat too much. Allow the bulb to cool; liquid enters to take the place of the expelled air. Repeat this process until the bulb and one inch of the tube is filled. When the bulb has been filled, place it in a pot of cold water and measure the height of the liquid on a paper scale attached to it, as in Expt. 6. Gently heat the water-bath, stirring occasionally. Observe the height of liquid every two minutes. Repeat this experiment, filling the bulb successively with alcohol and mercury, and verify the following statements:

1. That alcohol expands more, and mercury less, than water.
2. That mercury expands regularly, water and alcohol irregularly.

Allow them to cool, and observe how the same facts stand out prominently.

When we consider the case of water flowing from place to place, we always find that its flow is from a high level to a lower one—*i.e.*, water flows downhill, never uphill. In a similar manner, when a body feels hot, it is because the heat passes from the object to

the colder hand. Similarly, when heat passes from the hand to an object, the hand is cooled—*i.e.*, the object feels cold. It would seem, therefore, that heat can be, like water, sometimes at a higher “level,” we might say, than at others. We do not speak of heat-level as we do of water-level, but we use a special word, **temperature**, and we say that the **temperature of a body is that which determines the direction in which heat is transferred**—*i.e.*, from places at a high temperature to places where the temperature is lower.

We have seen above that as a liquid gets hotter—*i.e.*, as its temperature rises—it expands. Which liquid of those above do you think is the best to use for measuring temperature? Why?

An instrument used to **measure temperature** is called a **thermometer**, and we will now proceed to make this well-known instrument.

EXPT. 10.—To construct a Mercury Thermometer.—After a suitable liquid has been chosen—for ordinary purposes mercury is selected (Why?)—it is necessary to prepare a convenient bulbed tube, as in Expt. 8. Before commencing to fill it as there described, the tube must be drawn out thus. Warm the tube evenly about 1 inch from the end, rotating it in the flame until it is quite soft; remove it, and draw the ends apart *steadily and slowly*, not rapidly or jerkily. It should now appear as shown (Fig. 6). The mercury to be used should be almost boiled and allowed to cool somewhat, but not to become cold. Now fill the bulb and the *whole of the stem*, as described in Expt. 9. We must now decide for what purpose we desire to use our thermometer, and it will serve our ends if our thermometer indicates the temperatures between the freezing and boiling points of water, and a little below

and above respectively. Since our thermometer is open to the air, dust may get in, mercury evaporate away, and so forth. To avoid these defects, the tube is closed in the following way: Place the thermometer in a tall tin or flask of water or brine and boil the liquid briskly. Observe the expulsion of globules of mercury. When no more mercury is expelled, gently warm the thinned out portion, and finally heat it with a small Bunsen flame until the glass melts and closes. Remove the extra portion and the thermometer from the hot water. If the end is not sealed strongly enough, allow the thermometer to cool, when the end may be thickened by remelting the tip in the Bunsen flame.

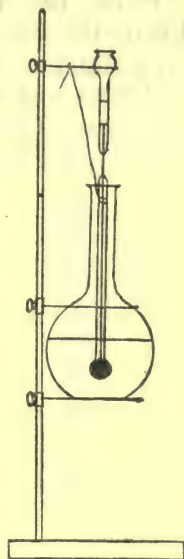


FIG. 6.

To graduate a Thermometer.—In order to graduate a thermometer, it is necessary to choose two temperatures which are always the same. Two are commonly used—*e.g.*, when we say “as cold as ice” or “boiling hot.” These temperatures are termed the **fixed points** of a thermometer.

EXPT. 11.—To determine the “Freezing-Point.”—Place your thermometer in a beaker, a quarter filled with water, and add to it cleaned crushed ice or snow, and observe the mercury every half-minute or so, stirring gently. The mercury falls until a certain point is reached, and then remains stationary for a long time—indeed, until all the ice has melted. Draw a file gently across at this point, making a fine mark.

This is the freezing-point of water or the melting-point of ice.

EXPT. 12.—To determine the “Boiling-Point.”—Fit up the apparatus shown, which consists of a flask (a) containing distilled water. Into the neck of the flask fix a cork and wide tube about 9 inches in length (b). Now support the thermometer by means of a thin slip of cork, so that it hangs well in the neck of the flask. Boil the water in the flask (a) briskly, and observe the mercury. It continues to rise until it reaches a particular level, just as the water begins to boil, and remains stationary.

Mark this level as before. This is the “boiling-point” of water.

Having now obtained our fixed points, we must decide what numbers to give them. For ordinary household purposes the freezing-point is called 32° F.—i.e., 32 degrees Fahrenheit—and the boiling-point 212° F., the space between them being divided into 180 equal parts, called “degrees.” Such a thermometer is a Fahrenheit thermometer, and you will no doubt see one of these hanging in the room.

For scientific purposes a simpler graduation is employed. The lower fixed point is labelled 0° C. (nought degrees centigrade), the upper one 100° C.; whilst the space between these is divided into 100 equal parts; hence its name *centigrade*—100 degrees. This division can be carried below the freezing-point, read as -15° C., and also above the boiling-point—e.g., 120° C. Often-times the scale is printed on paper which is attached

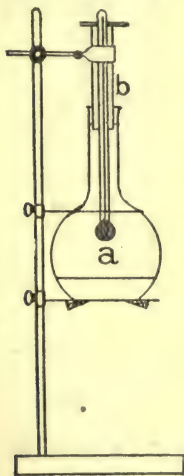


FIG. 7.

to the tube, the whole then being enclosed in a wider glass tube; but more often, especially in the case of chemical thermometers, it is etched on the glass itself.

We are now in a position to discover which of two bodies is at the higher temperature—*i.e.*, at a higher heat “level”—and we can, by the aid of a thermometer, carry our investigations a little farther. Although we can truly say that temperature corresponds to water-level, we have yet to discover how to measure the actual *quantity of heat* a body possesses. Let us make a few simple experiments with water first.

EXPT. 13.—To study the Flow of Water at Different Levels. — The vessel *A* (Fig. 8) contains a little water, and, resting on the surface, is a wooden float, carrying a 10-gm. weight. It is connected, by means of the narrow india-rubber tubing, carrying a clip or pinch-cock at *C*, to the

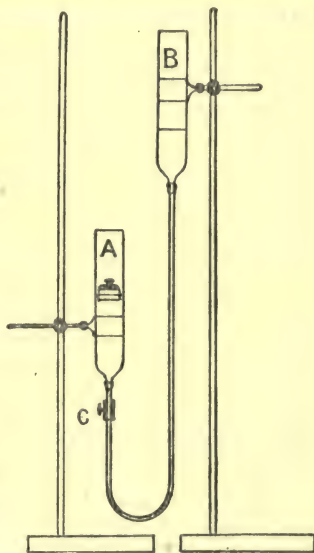


FIG. 8.

similar vessel, *B*, containing 25 gms. of water. Elevate *B* until the water-level in *B* is 1 foot above the weight in *A*. Open *C*, and observe the time taken by the water in *B* to run into *A*, and measure the rise of the float and weight. Repeat this experiment, making the water-level in *B* 2, 3, and 4 feet above the weight

successively, and verify the following statements from your results:

1. The rate of flow depends upon the height of the water-level in *B*.
2. The rise of level in *A* is independent of the height of *B*.

EXPT. 14.—To observe the Behaviour of Different Quantities of Water at the Same Level.—Repeat Expt. 13, placing 50 gms., 75 gms., and 100 gms. of water in *B* successively, and keep the level in *B* in each case 1 foot above the weight. Does not your experiment prove that the rise of level in *A* depends upon the quantity of water in *B*?

EXPT. 15.—To compare the Behaviour of Different Liquids.—Repeat Expt. 13, using 25 gms. of water, mercury, and alcohol successively in *B*, and verify the following statement: The rise of level in *A* depends upon the nature of the liquid.

Let us now perform similar experiments with regard to heat, and compare our results.

EXPT. 16.—To compare the Behaviour of Water at Different Temperatures.—Place 25 gms. of water in a clean beaker, and observe its temperature (15° C.). Now place another 25 gms. of water in a separate beaker and warm until its temperature is 35° C. Pour the warm water into the cold water, stir, and observe its temperature; it is 25° C.—*i.e.*, the temperature has risen 10° C. Observe the time taken to reach the constant temperature. Repeat this experiment, using 25 gms. of water heated to 75° C., and again observe the time taken, and the final temperature, 45° C.—*i.e.*, the temperature has risen 30° C. By comparing the times, do we not find that the heat necessary to warm the cold water up 30° C. is not three times, but much

less than the time required to raise its temperature 10° C.—*i.e.*, the higher the temperature, the more rapid is the “flow” of heat. Compare this conclusion with that obtained in Expt. 13 (1). Do we not also learn that the higher the temperature, the greater is the quantity of heat given out? Notice how different this result is from that obtained in Expt. 13 (2).

EXPT. 17.—To compare the Behaviour of Different Quantities of Water at the Same Temperature.—Repeat Expt. 16, using (1) 25 gms. of water at 75° C.; (2) 50 gms. of water at 75° C. In the former case the final temperature is 45° C.—a rise of 30° —and in the second it is 55° C.—a rise of 40° —so that the heat given out depends upon the quantity of materials.

EXPT. 18.—To observe the Behaviour of Different Materials.—Repeat Expt. 16, using (1) 50 gms. of water at 75° C., and (2) 50 gms. of mercury at the same temperature. In the former case the final temperature is 55° C., whilst in the latter it is only 19° C.—*i.e.*, a rise in the case of water of 40° C., in the case of mercury only 4° C.—so that we learn that the quantity of heat depends upon the nature of the substance.

From these experiments we may thus conclude—

1. The rate of transfer of heat depends upon the temperature.

2. The quantity of heat depends upon (1) the nature of the substance; (2) the quantity of the substance; and (3) temperature; and we therefore define the unit quantity of heat (a calorie) as that amount which will raise the temperature of 1 gm. of water one degree centigrade.

For methods of determining quantity of heat, the student should consult some textbook on Heat.

CHAPTER II

THE HISTORY OF CHEMISTRY

The Early Growth of Chemistry.—The matter surrounding us and its changes have always been the subject of interest and curiosity from the earliest times. Many isolated facts, accidentally obtained, were known and utilized by the Egyptian priests in the “holy arts.” This knowledge spread to the Phœnicians and Jews (the latter knew of gold—from Ophir; silver, copper—from Cyprus; iron, tin, etc.); and later to the Greeks and Romans by the writings of Solon, Pythagoras, Democritus, and Plato. The Greek philosophers speculated on the constitution of bodies without any regard for facts, and their “doctrine of the elements,” originating with Empedocles (440 B.C.), and developed by Aristotle, stamped the earliest ages, and held sway until medieval times. According to this doctrine, the basis of the world consists of four “principles”—**fire, air, earth, and water.**

The early goldsmiths imitated gold and jewels (by coloured glass, etc.), and had tests for detecting frauds by heat; and it is possible that such practices, supported by many *merely superficial observations*—e.g., the deposition of copper on iron tools in the copper-mines, alloys of copper and arsenic being white (hence called “silver”), alloys of copper and zinc being yellow (so-called “gold”), gave rise to the belief in the **transmutation of metals**, the mere change in colour being regarded as sufficient evidence of the change.

Alchemy.—When, in 640, the Arabians overran Egypt, they fostered the Egyptian arts (known as “*chemia*”), giving to the word the Arabic form *Alchemy*. Many new operations were discovered and new apparatus devised—e.g., the water-bath, improved furnaces, sublimation, filtration, distillation,

crystallization, etc. The main aim of the alchemists was to find the **Philosopher's Stone** of miraculous power, which would convert all common or base metals into the noble metals, gold and silver. In this search facts accumulated rapidly, but the deductions, based as they were on merely **qualitative** information, led to many fanciful beliefs and erroneous ideas. From the Arabians the study spread through Spain to Western Europe, one of the most famous of the European alchemists being **Basil Valentine**, c. 1450. He was a great practical chemist; his study of antimony and its compounds was a work of great completeness, and he advised the use of chemicals as medicines. Probably owing to this the idea arose that the Philosopher's Stone was also a universal medicine—the **elixir vitæ**—and even the creator of living things. The first tendency to **classification** arose about this time—*e.g.*, all soluble bodies were designated *sal* (*e.g.*, *sal vegetabilis* = carbonate of potash); whilst all volatile substances were termed *spiritus* (*e.g.*, *spiritus salis* = hydrochloric acid, still commonly called “spirits of salt”).

From the time of the Reformation (c. 1550) chemistry became somewhat separated from alchemy, although transmutation was still firmly believed in. It was now regarded as the basis of medicine, and a century of **preparations** began. **Paracelsus** (1493-1541)—the father of medicine—joined chemistry and medicine, and life processes were considered as chemical changes dependent on the animal organs and juices. **Van Helmont** (1577-1644) first grasped the idea of the “law of conservation of matter,” and was the real founder of “pneumatic chemistry” by his study of **gas sylvestre** (carbon dioxide), which he obtained from limestone and potashes by the action of acids, from coal and fermentation. He also described two combustible gases (hydrogen and marsh gas). **Glauber** (1604-1668), from his experiments with sulphuric acid on nitre and salt, gave us the idea of chemical *liking* or *affinity*, and our first examples of *double decomposition*. Throughout this period the fanciful alchemical ideas were still dominant, and obscured the real interpretation of the results obtained, and although the arts were enriched, the science itself was but slightly benefited. A change, however, was slowly coming, and out of the “clouds of obscurity” the new science was slowly appearing.

Matter is continually undergoing Changes.—When winter comes, the water in ponds, etc., becomes frozen into a glassy solid ice, which melts and forms liquid water again as the weather becomes warmer. How different is this change from the changes occurring when a piece of bright iron is left in the moist air, and a brownish powder—**iron rust**—is formed, until all the iron has disappeared, never to return; when milk turns sour, and acquires strikingly different properties as a result of exposure to the air; when coal, wood, and numerous other bodies—known as *combustibles*—burn on a light being applied, leaving in most cases no visible residue except a small quantity of “ash”! The former are known as **physical changes**, the latter as **chemical changes** or **chemical actions**. Let us study, then, these changes in more detail.

PHYSICAL CHANGES.

EXPT. 19.—To study the Action of Heat on Platinum.—Weigh a piece of clean platinum foil and heat it strongly in the Bunsen flame; it glows brightly. Remove it from the flame, and observe that it returns to its original state. When cool, reweigh; its weight is unaltered.

EXPT. 20.—To study the Effect of rubbing Steel with a Magnet.—Weigh a steel knitting-needle and rub it six times from end to end with a magnet. Examine it. Is it steel? Reweigh it; its weight is unchanged. Place in it some iron filings or steel pins; they are attracted. It has the new property of being *magnetic*, but it is still steel. Stroke it again in the opposite direction the same number of times, or hit it smartly with a hammer, and test with iron filings; it has lost

its magnetic property, and is once more the original steel knitting-needle.

Thus we see that when a body undergoes a **physical change** it acquires new properties, usually **temporary**, but **throughout the change** (before, during, and after) it consists of the **same kind of matter**—*i.e.*, no new **substances** are formed—and the **same quantity of matter**—*i.e.*, nothing is added to it, nothing taken from it. It is the original body, with new properties for the time being, and it can easily be changed back into its former state.

CHEMICAL CHANGES.

EXPT. 21.—To study the Action of Heat on Magnesium.—Place some magnesium ribbon in a crucible with a lid, and weigh the whole. Now heat this strongly on a pipe-clay triangle on a tripod stand, raising the lid occasionally. The magnesium burns with a bright white flame, and a greyish-white *earthy* powder—impure magnesia [A]¹—results, absolutely different in appearance and properties from magnesium. Cool and reweigh; a gain in weight is found.

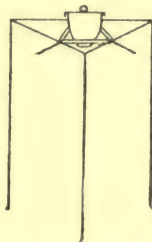


FIG. 9.

EXPT. 22.—To study the burning of a Match.—Weigh a dozen matches and strike them on the box. They burst into flame. Extinguish them and reweigh; a loss in weight is found. Examine the residue; an entirely different substance remains, and no amount of rubbing will make them burn again.

These experiments show us that when a body

¹ [] Where these references occur, turn to Chapter XXVI., p. 323, for their respective equations.

undergoes a **chemical change**, **totally new substances are produced**. They no longer consist of the original kind of matter only, but **usually something has been added to it or taken from it**; nor can the original body be regained.

We may thus define a **physical change** as a change in which no new kind of matter is formed at any time of the change, and a **chemical change** as a change which results in the formation of entirely new kinds of matter.

EXERCISE 1.—Find what kind of change occurs in each of the following experiments: (1) When sugar is heated; (2) when glass is heated; (3) when salt is dissolved in water; (4) when zinc is dissolved in dilute sulphuric acid; (5) when ammonium chloride is heated; (6) when lead nitrate is heated; (7) when “soil” is heated.

It is the business of the chemist to investigate and find out, as far as he can, the nature of the substances taking part in chemical actions, the causes and results. **Chemistry** may, then, be defined as **that branch of science (knowledge) which treats of the composition of matter and of the changes which matter undergoes**. It is an *experimental* science—i.e., the student must put the question to Nature, by means of an experiment, with his own mind and hands, and closely observe her answer.

It is obvious that, before any investigation of the nature of a substance can be undertaken, it is necessary to obtain that substance in a pure state. We therefore proceed next to the methods used by the chemist to prepare *pure* substances.

SOLUTION.—When a solid placed in a liquid disappears and is afterwards found in every part of the liquid (by taste, evaporation, or other means)—e.g., sugar in tea or salt in water, etc.—it is said to **dissolve**

(or to be **soluble**), and the resulting liquid is termed a **solution**. The liquid which dissolves the solid is termed a **solvent**.

EXPT. 23.—To separate a Solid from a Solution (Evaporation).—Place some common salt in a dish half full of water, and stir until all has dissolved. Now place the dish and solution on a piece of wire gauze on a tripod stand, and gently heat with the Bunsen flame. As the water boils away in the form of steam or vapour—*i.e.*, *evaporates*—crystals of salt will be found. What kind of change is this ?

Definition.—**Evaporation** is the conversion of a liquid into gas or vapour.

Now, not all solids will dissolve in water—*e.g.*, repeat Expt. 23, using a piece of marble, and evaporate the clear liquid. This leaves no deposit, showing that marble is a solid which is *insoluble in water*. Hence we might use the different behaviour of two solids with respect to water as a means of separating them.

EXPT. 24.—To separate a Mixture of Two Solids, only One of which is soluble in Water—1. Decantation.—Take a quantity of a mixture of sand and salt. Place this in a test-tube or beaker, add water, shake well, and allow the sand which is in *suspension* to settle. Carefully pour off the clean solution on the top, without disturbing the sediment, into an evaporating basin. This process is known as *decantation*. Evaporate the clean decanted solution to dryness, and crystals of salt appear. Taste it, and note that it is pure. The sand may be washed (Why ?) and dried if required.

Definition.—**Decantation** is the gentle pouring of a liquid from a sediment, or from one vessel into another without disturbance.

For the next process of separation a *stirring-rod* is required. This may be solid or hollow, and can be made in the following way:

To make a Solid Stirring-Rod.—Cut off a piece of glass rod, as in Expt. 3, about 5 or 6 inches long, and round the ends, as described in Expt. 4.

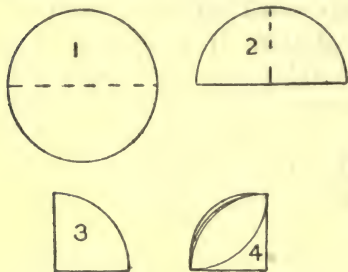


FIG. 10.

To make a Hollow Stirring-Rod.—Cut off the same length of medium-sized glass tubing, and round off each end until each is just closed. Which of these stirring-rods would you use for glass apparatus? Why?

2. Filtration.—Repeat the above experiment, using a mixture of chalk and salt. In this case the chalk is so finely divided that it remains in suspension, giving the liquid a milky or muddy appearance, and decantation is not suitable. Under these circumstances *filtration* is resorted to. Take a filter paper and fold it *accurately* twice, as shown (Fig. 10): open it into a cone, place in a glass funnel which it will just fit, and moisten with pure water. Grease the lip of the beaker or test-tube with *vaseline*. (Why?) Pour the milky liquid down a glass rod into the filter (Fig. 11), *taking care never to fill the filter more than two-thirds full*. (Why?) When all has been transferred, the suspended chalk remains on the filter-paper, whilst the clear liquid which runs through—the



FIG. 11.

filtrate—contains the salt, which is obtained by evaporation. As before, by washing (*let each wash-water completely drain out before adding the next; why?*) and drying, the chalk may also be obtained pure. *The stirring-rod must never be placed on the bench. (Why?) Where should it be kept?*

Definition.—Filtration is the process of separating a liquid from an insoluble solid by passing through some porous material.

Compare decantation with filtration. If the substances are both insoluble in water, then a suitable solvent must be found for one of them.

EXPT. 25.—To find whether a Substance is Soluble or Insoluble in Water.—Mix a little of the substance with water in a test-tube; shake well, and warm. Decant or filter off the clear liquid from any residual solid, and evaporate to dryness. If any solid remains, then the substance is soluble; if not, it is insoluble.

EXERCISE 2.—Find out which of the following are soluble and which are insoluble in water: Sand, chalk, tin, sal ammoniac, blue vitriol, alum, nitre, plaster of Paris, charcoal, washing-soda, green vitriol, sugar, wood, sulphur, clay.

EXPT. 26.—To find a Solvent for Sulphur.—Since sulphur is insoluble in water, before we can use the above method of separation we must find a liquid in which it does dissolve. Repeat Expt. 25 with sulphur, *powdered* (Why?), and using first methylated spirit, then carbon disulphide.

Caution.—*Do not warm or evaporate these liquids over the Bunsen flame, since they are very inflammable, but allow them to stand on the bench or on the top of the steam oven until all the liquid has disappeared. No residue is found in the former case, but sulphur remains in the second.*

Thus carbon disulphide is a solvent for sulphur.

EXERCISE 3.—Find a solvent for (1) shellac; (2) bees-wax; (3) charcoal; (4) iron; (5) chalk. Try water, methylated spirit, carbon disulphide, dilute sulphuric acid, dilute hydrochloric acid. What kind of change occurs in each case?

EXERCISE 4.—Using the information you have gained in Exercises 2, 3, separate the given mixtures: (1) Sulphur and chalk; (2) plaster of Paris and common salt; (3) sulphur and shellac; (4) gunpowder (charcoal, sulphur, and nitre); (5) soil (sand, clay, etc.).

CRYSTALLIZATION.—When two substances are mixed together, both of which are soluble in a liquid, but in different degrees, another plan must be adopted.

EXPT. 27.—To obtain Crystals of Nitre (Potassium Nitrate).—Half fill a test-tube with water and add to it *powdered* (Why?) nitre in small quantities, shaking after each addition, until no more will dissolve—*i.e.*, a little remains undissolved after continued shaking. The solution is said to be **saturated**. Decant the clear liquid, set it aside and examine later. Repeat the above, but warm the liquid. Observe that a **much larger quantity will dissolve in hot water than in cold**. Cool rapidly by placing under the tap and allowing cold water to run over the outside. Compare the crystals obtained in both cases. What kind of cold solution is left in the second case?

We observe that the first process gives us *few large* crystals, whilst the latter gives *many small* crystals. Which are the purer?

Crystals are thus obtained (1) from a weak solution by the evaporation of the solvent until too little remains to dissolve the solid; (2) from a strong solution

by cooling. They are obtained in other ways, which we shall meet later. We may define **crystals**, then, as bodies which in forming take a definite geometrical shape, bounded by plane faces; whilst **crystallization** is the process of forming crystals.

In order to grow crystals of a larger size and more perfect kind, the following methods may be used:

EXPT. 28.—To make a Large Crystal of Bluestone (Copper Sulphate).—Powder some bluestone in a mortar, and add, in small quantities at a time, to half a beakerful of water, stirring until some remains undissolved, and set it aside until the next lesson. From the crystals formed pick out the best in shape (and size), and proceed thus: (1) Tie this selected crystal by a thread or horsehair to a wooden spill and suspend it in the solution, replenishing it as required by a saturated solution, prepared as before; or (2) place the crystal in a clean beaker on the side which was uppermost (Why?), and pour over it the solution. Repeat this process every day until the crystal is of the required size. Sketch the shape of the large crystal you have prepared.

EXERCISE 5.—Prepare large and small crystals of alum, sugar, sulphur, potassium bichromate.

EXPT. 29.—To separate Potassium Chlorate from Potassium Chloride.—Mix together nine parts of potassium chlorate with one part of potassium chloride. Warm half a beakerful of water, and add the mixture in small quantities, until some remains undissolved. Decant the clear liquid into another beaker and allow it to cool. When it has crystallized, pour off the clear liquid. Wash the crystals with a *small quantity of water* (Why?), and dry between filter-papers. Test

them thus: Into three clean test-tubes place a few c.cms. of distilled water, and to the first add one crystal of potassium chloride; to the second, one crystal of potassium chlorate; and to the third, one of the crystals you have prepared. Add to each one of these a spot or two of silver nitrate solution. We observe that potassium chloride gives a white solid (**precipitate** of silver chloride) [**B**], but potassium chlorate does not. What does your preparation give? Is it potassium chlorate? Where is the potassium chloride? Test and verify your conclusions.

If your substance is not quite pure, repeat the process (**recrystallization**) until you obtain pure potassium chlorate.

EXERCISE 6.—Prepare some pure washing-soda and bluestone from the impure samples supplied. Obtain some pure nitre and sulphur from gunpowder.

DISTILLATION.—In the experiment on evaporation we only concerned ourselves with the solids that remained, but in many cases the liquid is also required. In the laboratory, where pure water is essential, it is obtained by the process of **distillation**—i.e., by evaporation and condensation—or by converting a liquid into vapour and condensing the vapour back into liquid in a different vessel.

EXPT. 30.—To prepare **Pure Water from Impure Water by Distillation.**—Place in a retort (*a*) some water made impure by the addition of a few crystals of bluestone. Fit it up on the retort stand, over a wire gauze, as shown (Fig. 12), placing the neck (tubulure) of the retort into a flask (*b*) floating on cold water in the trough (*c*). Now boil the liquid in the retort, neglecting the first portion which distils over (Why?), and collecting about two-thirds of the remainder. (Why no more?) Now

compare the liquid residue in *a* and the distillate in the receiver (*b*) as to (1) colour, and (2) their behaviour

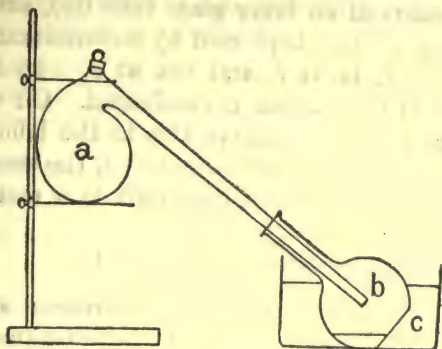


FIG. 12.

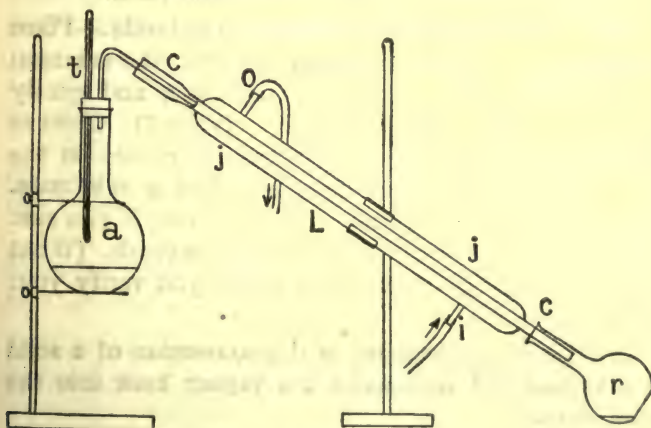


FIG. 13.

on evaporation. In this simple apparatus a considerable amount of steam is lost. (Suggest how to prevent this ?) The apparatus shown in Fig. 13 is used to

avoid this. A distilling-flask (*a*) is fitted with a thermometer (*t*), and attached to a Liebig's condenser (*L*), which consists of an inner glass tube (*cc*), surrounded by a "jacket" (*jj*), kept cold by a continuous stream of cold water, in at *i*, and out at *o*. In this way the whole of the vapour is condensed. On the large scale—and you may observe this in the laboratory—a copper boiler is usually employed, the steam being condensed in a tin worm immersed in a metal vessel of cold water.

What advantages do these possess ?

SUBLIMATION.—This is a process somewhat similar to distillation, save that the substance is converted into vapour, and, by cooling, the vapour is condensed in the **solid**, instead of the liquid form.

EXPT. 31.—To sublime Ammonium Chloride.—Place a crystal or two of ammonium chloride (the common sal ammoniac) in a small dry test-tube, and gently warm the end of the tube only. (Why ?) Observe that no melting occurs, whilst crystals collect on the cool sides of the tube. This is called a **sublimate**. Examine these with a lens, and sketch one, if you can. This is, of course, a method of making crystals. (What kind of change is this ? How would you verify your answer ?)

Definition.—Sublimation is the conversion of a solid into vapour and condensing the vapour back into the solid state.

Bodies which, like ammonium chloride or water, *sublime* or *distil*, are said to be **volatile**; whilst bodies like sand or bluestone are termed **fixed**, or **non-volatile**.

EXPT. 32.—To separate a Volatile Substance from a Non-Volatile Solid in a Mixture of Both.—Gently heat a

little of the mixture of sand and sal ammoniac supplied, and observe that the ammonium chloride collects as a white **sublimate** on the sides of the tube, whilst the non-volatile sand remains pure at the lower end.

EXERCISE 7.—Examine the action of heat on sulphur and salt, and separate a mixture of the two. (Could they be separated in any other way? Suggest methods.)

IGNITION.—When bodies are ignited—*i.e.*, **intensely heated**—they may or may not burn—*e.g.*, heat a piece of lime, asbestos, platinum, etc. Occasionally the chemist avails himself of the combustibility of certain bodies, and burns them off from others which are unalterable by heat.

EXPT. 33.—**To separate Sand from Sulphur.**—Place some of the mixture supplied in a crucible and ignite strongly; the sulphur burns away, leaving pure sand.

EXERCISE 8.—Obtain some pure lime from the mixture of lime and charcoal, some pure salt from starch and salt.

Being now in a position to prepare different kinds of matter in a pure state, let us see if all matter is of one type or not.

As in the following experiments it will be necessary to fit up special apparatus, we proceed first to learn how to bend glass tubing to any required shape.

EXPT. 34.—**To bend a Glass Tube.**—Cut a convenient length (about 5 inches) of suitable tubing, and, after warming, hold this in the upper yellow portion of the bat's-wing or fish-tail flame, *parallel with the length, and not across it*, rotating it until it begins to "sag" or "bend," and then hold it perfectly still in the flame until it falls to the required angle. Remove it, and allow it to cool.

Caution.—Never try to force glass while it is in the flame. You may remove it from the flame and bend it between the fingers, but the method described above is the better one. Now remove the soot with a *dry* duster and round the ends.

TYPES OF MATTER.

EXPT. 35.—To find out the Kind of Matter in **Mercuric Oxide**.—Gently heat a little of the red powder, mercuric oxide, in a small tube and observe the appearance of the sublimate. It consists of little beads

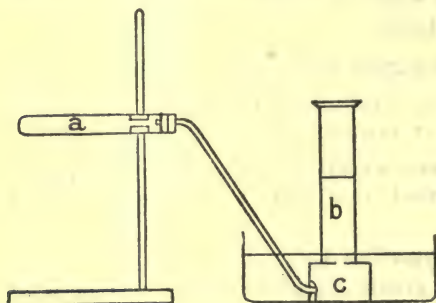


FIG. 14.

of *quicksilver* (or mercury). Plunge a red-hot piece of wood into the mouth of the tube; it bursts into flame. (Why?) Fit up the apparatus shown, and see if any gas can be collected.

While quicksilver remains in the tube (*a*), a colourless gas collects in the gas-jar (*b*), which relights a glowing piece of wood. This gas is called "*oxygen*." (What other property of this gas do you learn?) Thus, mercuric oxide has yielded two different kinds of matter—a liquid (*mercury*), and a colourless gas (*oxygen*) [*C*].

EXPT. 36.—To investigate the Type of Matter in **Lead Nitrate**.—Repeat Expt. 35, using the white, crystalline, solid lead nitrate. At first a *brown* gas comes off, and on plunging in a glowing wood ember,

this is relighted, whilst a *yellow solid* is left. Collect any gas you can over water and test it. It is *colourless, insoluble, and relights a glowing taper*. (Where has the brown gas gone?) Test the water in the trough with blue litmus—this is reddened—which is used as a test for acids. Thus, lead nitrate has yielded (1) a **yellow solid** (named **lead oxide**); (2) a **brownish gas**, which is **soluble**, and forms an **acid**; (3) a **colourless gas** (**oxygen**) [D].

Hence some kinds of matter are complex—*i.e.*, are made up of simpler forms. Such a chemical change as this, **when a complex body is split up into simpler ones**, is called **chemical decomposition**. Now, whatever we do to the liquid metal mercury or the colourless gas oxygen, we cannot decompose them into simpler forms. Whenever we produce new types of matter from these, we always find that the weight increases—*i.e.*, new substances can *only* be made by adding other kinds of matter to them. Such substances are called **elements** (or ground-stuffs), and we define an **element to be a substance which has never been split up into anything simpler**. These are the bricks and mortar of the chemist. Some eighty elements have been obtained, and all other forms of complex matter are obtained from these. We give below a list of a few of the more common:

Non-Metals.—Hydrogen, oxygen, nitrogen, carbon, sulphur, phosphorus, chlorine.

Metals.—Magnesium, mercury, lead, copper, iron, zinc, platinum, tin, aluminium, sodium, potassium.

The above experiments also teach us that **heat simplifies complex matter**.

The next question which arises is—Are the complex forms of matter all of the same type? In the preceding sections we saw that various kinds of bodies could be separated by solution, distillation, crystallization, sublimation, and ignition.

EXPT. 37.—To examine the Effect of Solution on Lead Nitrate.—Take a crystal or two of lead nitrate (used in Expt. 36) and shake it up with water. Observe that it *all* dissolves, not a part of it, as in Expt. 24. Evaporate the solution and examine the residue. Heat a little in a tube; it is the same substance after solution as it was before, so that solution does not alter some complex substances, whilst it does others. Distillation, crystallization, and sublimation, lead to the same result, effecting a separation of certain kinds of matter, but not all; so that **COMPLEX MATTER** may be of two types.

EXPT. 38.—To compare a Mixture with a Compound.—Mix some fine iron filings and powdered sulphur in a mortar; touch the mixture—it is quite cold. Divide this material into two portions, and place one aside, but heat the other in a dry tube. When the end of the tube is red-hot and the action has commenced, remove it from the flame and observe how the action continues of its own accord, **giving out light and heat**. When cool, break the tube, powder the substance (known as **iron sulphide**) [E], and proceed to examine it as below:

1. Observe the mixture and this substance with a pocket-lens or microscope. In the former, both iron and sulphur can be seen, but neither is visible in the latter.

2. Shake both up in separate tubes with water, and

allow to stand. In the former we have two layers visible—viz., of iron (lower) and sulphur. In the latter no such separation is effected.

3. Bring a magnet near to both, and observe that iron is withdrawn from the former, while sulphur is left, but not from the latter.

4. Shake up portions of each with carbon disulphide, decanting the clear liquids into clean dishes and allowing them to evaporate. Observe that while the former yields a residue of iron and its liquid leaves sulphur, the latter does not give either.

Thus, complex matter may be of two types:

1. **Mixtures**, like salt and sand, or iron and sulphur, of two or more bodies without union, in which each component still retains its own properties, and, because of this, their separation can be effected easily by physical means.; and

2. **Compounds**, like mercuric oxide or iron sulphide, of two or more bodies chemically united, in which each component loses its characteristic properties, and thus their separation cannot be effected by physical means.

Further, we see that, whereas the properties of a mixture are merely the properties of its components, those of a compound are quite different.

When two bodies unite or combine (as iron and sulphur above, or as the magnesium and oxygen in Expt. 21) to form an entirely new substance, a compound, it is termed "chemical union" or "combination." Not only does chemical union result in the formation of new bodies, but it is generally accompanied by an evolution of heat. These are two of the principal characteristics by which a chemical union is detected.

Before carrying out the following exercises, read the instructions for practical work on pp. 7, 8.

EXERCISE 9.—Try the effect of water, dilute sulphuric acid, and heat, on the following, and decide which are mixtures, compounds, or elements: Soil, mortar, chalk, soap, sugar, brass, tin, iron, galena.

EXERCISE 10.—Carefully study the action of heat, (a) gentle, (b) strong, on the following substances, noting down all that you observe **at once**, and collecting over water any gases you can. Examine the gases (1) with a light, (2) a glowing wood splinter, (3) lime-water. Also examine the solid residues if any, and the liquid in the trough with litmus: Sand, soil, ammonium bichromate, mercuric oxide, zinc carbonate, sulphur, washing-soda, alum, potassium chlorate, sodium formate, nitre, sal ammoniac, bluestone, green vitriol, wood.

CHAPTER III

SOLUTION

EXPT. 39.—To find whether any Change in Weight occurs during the Solution—

1. Of Salt in Water.—Into a beaker place about 40 c.cms. of water and a stirring-rod. Place a tea-spoonful of common salt in an evaporating-dish, or watch-glass and rest it in the mouth of the beaker, and weigh the whole. Now empty the salt into the water (What precautions must be taken?); stir until all has dissolved; replace the dish and reweigh. **Solution of salt in water is a physical change** (see Exercise 1, 3) and this, we now learn, **does not cause any change in weight.**

2. Of Zinc in Sulphuric Acid (Dilute).—Repeat the above experiment, but with these substances. This solution, we discovered, was a **chemical change**, and **a loss in weight occurs.**

Caution.—*The stirring-rod must not be taken out of the beaker; neither must the dish or rod be placed on the bench. (Why?)*

EXERCISE 11.—Repeat the above experiment, using nitre and water, bluestone and water, sulphur and carbon disulphide, iron and hydrochloric acid (dilute), copper and nitric acid (dilute). What conclusions do you arrive at?

MEASUREMENT OF VOLUMES.

In the section on the Balance we mentioned that the system of weights and measures adopted in scientific work was the metric system. The **unit of volume** for small measurements is the **cubic centimetre** (written **c.c.** or **c.cm.**); for larger volumes the **litre (l.)** is employed.

NOTE.—1 c.cm. of water at 4° C. weighs 1 gm.
 1 litre „ „ „ 1 Kgm.
 i.e., 1 litre = 1,000 c.cms.

28.4 c.cms. = 1 fluid ounce (approx.).
 1 litre = 1.76 pints (roughly $1\frac{3}{4}$ pints).

The instruments used in which volumes are determined are of two kinds—(1) those used to *measure* (or *hold*) a known volume of a liquid; (2) those used to *deliver* (or *pour out*) a known volume of a liquid. The commoner instruments of Class 1 are the graduated cylinder and the standard flask.

The Graduated Cylinder (i.e., a divided or marked cylinder) consists, as is shown in the sketch (Fig. 15, A), of a glass cylinder, marked with horizontal divisions denoting equal volumes. It is usually graduated into c.cms., each ten being labelled. Since this is used to **measure** a liquid poured into it, the instrument itself must be *dry* before use. (Why?) Pour 100 c.cms. of water first into a cylinder dry, and then into the same cylinder when wet, and compare. If you carefully examine the level of the liquid, you will notice *three* levels, and, as the diagram shows, it is of great importance to read the **lowest level of the liquid with the eye level with it**, or the reading will be in-

accurate. Thus, the correct reading given by the eye in Position 1 is 36 c.cms., whilst the reading given by an eye at 2 is too low—viz., 33 c.cms.—and that of an eye at 3 is too high—viz., 38 c.cms.

The **Standard Flask** (Fig. 15, B) is simply a flask with a long neck, and at a particular point on the neck

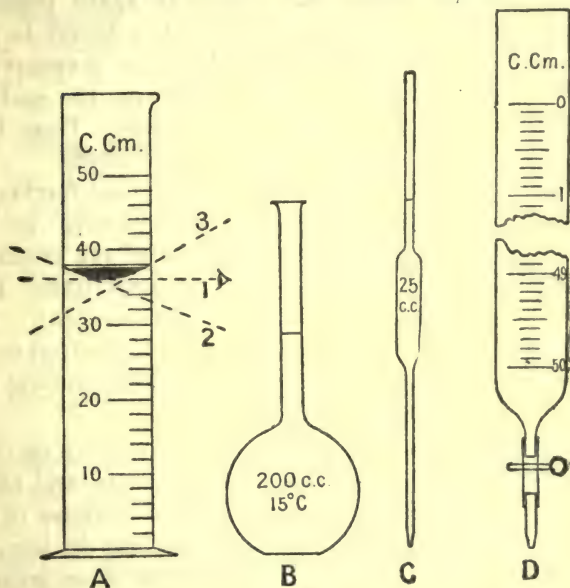


FIG. 15.

a mark is scratched. This flask *holds*, when filled to the mark, as shown, a definite volume of liquid, and before use must be quite dry. Why is the temperature marked on it?

The **Pipette** (Fig. 15, C) is a tube provided with a large cylindrical bulb in the middle, the lower end being drawn out to a jet. It has a mark above the

bulb, and is filled to this mark by suction. When the liquid is above the mark, the *dry forefinger* is quickly placed on the top, and the pipette lifted from beneath the liquid. By gently twisting the pipette with the other hand, the pressure on the top of the pipette is gently released, the liquid slowly falling. When it has reached the mark, the finger is again pressed tightly, and the pipette containing the liquid is removed to the vessel to receive it, the finger is removed, and the liquid allowed to drain out, with the end of the pipette touching the side of the vessel. Thus, the pipette **delivers a definite volume** of the liquid.

Caution.—When using a pipette for transferring a corrosive or poisonous liquid—*e.g.*, strong acids, alkalies, etc.—**the suction must be careful and gradual, and the end of the pipette must be kept under the liquid, or the liquid is drawn up into the mouth. In no case must the liquid be swallowed.** It is often convenient to attach a length of india-rubber tubing to the mouth of the pipette in such cases.

The Burette (Fig. 15, D) may be regarded as a special form of pipette, provided with a tap, or tube and clip, and is used for delivering any desired volume of a liquid. Note the following points regarding its use:

(1) The burette must be *clean and free from grease*, etc.

(2) Always use a funnel for filling, but remove it before using the burette.

(3) The *tap must be filled* before the first reading—*i.e.*, no air-bubbles enclosed. (Why?)

(4) The burette must be kept *vertical*. (Why?)

(5) The liquid must *not be run out too quickly*. (Why?)

EXERCISES.—Using water only, practise filling a pipette and then verify the accuracy of your instruments,

as below: (1) Fill a *dry* graduated cylinder with your pipette, comparing the readings after each addition; (2) repeat, using the cylinder and burette; (3) fill the *dry* standard flask with the burette and pipette, and note the reading required to fill it up to the mark.

EXPT. 40.—To find the Weight of Solid dissolved in any volume (say 100 c.cms.) of a Given Solution.—When we evaporated a solution of salt (Expt. 23), we noticed that towards the end of the operation “spitting” occurred, and solid was thrown out of the dish. How can this be avoided?

1. Slow Method.—If a water-bath is at hand, use it, but if not, take a beaker or tin-can, nearly fill it with water, and place in it some scraps of filter-paper. Weigh a clean evaporating-dish that will rest conveniently on the top of the beaker or can, together with a suitable stirring-rod. By means of a pipette or burette, place in this dish 20 or 25 c.cms. of the given solution. Put the dish and its contents on the *water-bath*, and boil the water in it briskly, filling it from time to time as required. When the evaporation of the solution is complete, gently heat the dish on a sand-tray or wire gauze over a low flame (Why?), stirring continually. (Why?) Cool and weigh; reheat for five minutes, cool and reweigh, and repeat until *two consecutive weighings are the same*. What does this tell you? Calculate your results as below:

2. Quick Method.—Weigh an evaporating-dish and rod as before; place in it 20 or 25 c.cms. of the solution, and evaporate on a wire gauze or sand-tray. As the evaporation proceeds, lower the Bunsen flame and stir continually, to avoid loss by “spitting.” Cool and weigh; reheat, cool, and reweigh until *the weight is constant*. Calculate your results as before, and compare it with your result obtained above.

Constant weight of dish, rod, solid = 28.15 gms.

Weight of dish, rod alone = 27.61 „

Weight of solid = 0.54 gm.

Volume of solution = 20 c.cms.

So that 20 c.cms. of solution contain 0.54 gm. of solid;

∴ 1 c.cm. of solution contains $\frac{0.54}{20}$ gm. of solid,

and 100 c.cms. of solution contain

$$100 \times \frac{0.54}{20} \text{ gm. of solid} = 2.70 \text{ gms. of solid.}$$

Which is the more accurate method? When *must* Method 1 be used? Give reasons for your answers.

EXPT. 41.—To find the Weight of Dissolved Solid in any Weight (say 100 Gms.) of a Given Solution.—This is really the same experiment as Expt. 40 above, except that the dish, rod, *and solution* must also be weighed before evaporating.

RESULTS—

Weight of dish, rod, solution = 48.35 gms.

Weight of dish, rod, alone = 28.73 „

∴ Weight of solution = 19.62 „

Constant weight of dish, rod, solid = 29.25 gms.

Weight of dish, rod, alone = 28.73 „

∴ Weight of solid = 0.52 gm.

So that 19.62 gms. of solution contain 0.52 gm. of solid;

∴ 1 gm. of solution contains $\frac{0.52}{19.62}$ gm. of solid.

and 100 gms. of solution contain

$$100 \times \frac{0.52}{19.62} \text{ gm. of solid} = 2.65 \text{ gms. of solid.}$$

EXPT. 42.—To find the Weight of Solid dissolved by any Weight (say 100 Gms.) of Water (or any other Liquid) in a Given Solution.—When the liquid is an

inflammable one, the slow method of Expt. 40 must be used; but if not, either method may be employed, as above.

RESULTS—

Weight of dish, rod, solution = 48.35 gms.

Constant weight of dish, rod, solid = 29.25 „

∴ Weight of water = 19.10 gms.

Constant weight of dish, rod, solid = 29.25 gms.

Weight of dish, rod, alone = 28.73 „

∴ Weight of solid = 0.52 gm.

So that 19.10 gms. of water dissolve 0.52 gm. of solid;

∴ 1 gm. of water dissolves $\frac{0.52}{19.10}$ gm. of solid,

and 100 gms. of water dissolve

$$100 \times \frac{0.52}{19.10} \text{ gm. of solid} = 2.73 \text{ gms. of solid.}$$

EXPT. 43.—To determine the Amount of an Insoluble Solid in any Quantity (say 100 Gms.) of a Mixture containing it, together with a Soluble Substance.—In Expt. 24 we learned how to separate a soluble solid from an insoluble one when mixed together, and, further, that filtration was better than decantation for this purpose. We therefore proceed as below:

Take two filter-papers of the same size and adjust them until they are of the same weight by cutting off small pieces from one of them. Place on the *torn* filter-paper about a gram or two of the mixture supplied (sand and salt), and put this on the left-hand pan of your balance. On the other pan place the untorn filter-paper and weigh the mixture. Fold the untorn paper and pour the mixture into it; then, folding the torn one, place it outside the other. Place the whole in a funnel and pour small quantities of *hot* water from

a test-tube down a glass rod on the mixture, allowing each portion to drain completely before adding the next. (Why?) A solution of salt runs through, and should be collected in a clean beaker. Repeat the washing until all the salt is washed out. (How will you know?) If you evaporate the last spot to dryness, what will the result be if all the salt has been washed out? If not? We have now all the *soluble substance in the filtrate* and the *wet insoluble substance on the filter-paper*. Dry the latter in a steam oven or drying-cone (Fig. 16).



FIG. 16.

Caution.—Be careful in the use of the latter apparatus *not to heat too strongly* and *char* the papers. When dry, separate the two filter-papers and reweigh. Redry, cool, and reweigh until of constant weight.

If there is any doubt as to the accuracy of your result, this may be checked by evaporating the filtrate in a weighed dish with a rod, as in Expt. 40, and weighing the dried soluble solid.

RESULTS—

Weight of mixture = 2.06 gms.

Constant weight of insoluble = 0.74 gm.

\therefore Weight of soluble = 1.32 gms.

So that 2.06 gms. of the mixture contain 0.74 gm. of insoluble;

\therefore 1 gm. of the mixture contains $\frac{0.74}{2.06}$ gm. of insoluble, and 100 gms. of the mixture contain

$100 \times \frac{0.74}{2.06}$ gms. of insoluble = 36.39 gms. of insoluble,

and $100 - 36.39 = 63.61$ gms. of soluble,

EXERCISE 12.—Find the amount of chalk in 100 gms of the mixture of chalk and ammonium chloride; also the amount of nitre, sulphur, and charcoal in gunpowder, and of chalk in the mixture of chalk and sand provided. Read Experiments 25, 26 before commencing these.

To make a Saturated Solution.—In Expt. 27 we found that a certain quantity of a liquid would only dissolve a limited quantity of a solid, and also that a hot liquid would dissolve more than the same quantity of the same liquid when cold. Hence, to make a **saturated solution**—*i.e.*, one which cannot dissolve any more solid under ordinary conditions—we may proceed in two ways.

EXPT. 44.—**To make a Saturated Solution at the Ordinary Temperature :**

Method 1.—Place some powdered solid (Why ?)—*e.g.*, nitre or copper sulphate—in a test-tube or bottle, and add some cold distilled water. Shake vigorously, or stir continuously, or blow air through the liquid by means of a pair of bellows, etc., for five or ten minutes. If *all* the solid dissolves, add more of the substance, and repeat until *some solid remains undissolved*. Allow the mixture to settle, and decant or filter off the clear solution.

Method 2.—Add a fairly large quantity of the powdered solid to some distilled water in a test-tube or beaker, and warm until the whole is in solution. Cool this hot, strong solution under the tap, or by standing it in a vessel of cold water, and when *quite cold*, crystals should be thrown down—*i.e.*, *precipitated*—a cold saturated solution remaining. If crystals do not form, more solid must be added and the process repeated.

EXPT. 45.—To make a Saturated Solution at any Required Temperature (say 45°C.):

Method 1.—Place a beaker or tin containing tap-water on a wire gauze and gently heat until the temperature of the water is about 45°C. Now place in this a large test-tube containing water and the powdered solid, together with a thermometer. Stir the solid and water together, and, by adjusting the flame of your Bunsen, keep the temperature constant at 45°C. for ten or fifteen minutes. Allow the mixture to settle *in the bath* and *rapidly* decant the clear saturated solution at 45°C. (Why not filter?) Instead of stirring, air may be blown through, as before. What disadvantages do you discover in this method? Suggest how to avoid them.

Method 2.—Proceed as in Expt. 44 (2) above, making a hot strong solution at a temperature above 45°C. , and cool it in a bath kept at this temperature, and when the temperature is stationary at 45°C. , decant as above. Which of these experiments do you think the better? Why?

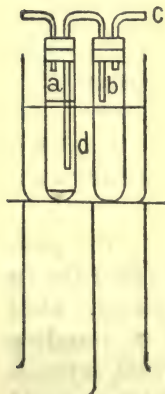


FIG. 17.

The chief disadvantage in the above method lies in the fact, that while decanting, the solution cools and crystals form. To avoid this, the experiment can be performed as follows: Two boiling tubes are fitted with corks and tubes, arranged as shown (Fig. 17); *a* contains the liquid, solid, and thermometer; *b* is empty and weighed. Both are placed in the bath *d*, and the mixture is stirred by blowing air through *c*. When the saturated solution is obtained, it is siphoned from

a to *b* by withdrawing air at *c*, and subsequent cooling is immaterial. This is especially important for accurate solubility experiments (see below).

EXERCISE 13.—Find the amount of solid in 100 c.cms. of a saturated solution at the ordinary temperature of each of the following: Chlorate of potash, bluestone, calcium chloride, washing-soda, common salt.

EXERCISE 14.—Find the amount of solid in 100 c.cms. of a saturated solution of salt in alcohol and of calcium chloride in alcohol, and compare your results with those of Exercise 13.

EXERCISE 15.—Compare the amounts of solid in a saturated solution of chlorate of potash at the ordinary temperature, 20° C., 40° C., 60° C., and of bluestone at the ordinary temperature, 35° C., 55° C., and 75° C.

From these experiments do we not learn that—

1. The amount of different solids that dissolve in a given quantity of the same liquid at the same temperature *depends upon the solid*.

2. The amount of the same solid that dissolves in the same quantity of different liquids at the same temperature *depends upon the liquid*.

3. The amount of the same solid that dissolves in a given quantity of the same liquid at different temperatures *depends upon the temperature*, and as a rule the higher the temperature, the greater the amount dissolved, other things being the same.

The “ease” of solution of a substance or its solubility is therefore defined to be the number of grams of a solid that will dissolve in 100 gms. of a liquid (usually water) at the required temperature, and this, we have learned above, depends upon—

1. The nature of the solid.
2. The nature of the liquid.
3. The temperature.

In order, then, to determine the solubility of a substance, we must *first* make a saturated solution of the given solid in the given liquid at the given temperature (Expt. 45); and, *secondly*, determine the amount of the solid dissolved by 100 gms. of the liquid in this solution, as in Expt. 42. The calculation of the solubility is then carried out as follows:

RESULTS.—In a saturated solution at 40° C.—

Weight of dish, rod, and saturated solution = 39.52 gms.

Constant weight of dish, rod, and solid = 30.27 „

∴ Weight of water = 9.25 „

Constant weight of dish, rod, and solid = 30.27 gms.

Weight of dish, rod, alone = 29.03 „

∴ Weight of solid = 1.24 „

So that 9.25 gms. of water dissolve 1.24 gms. of solid;

∴ 1 gm. of water dissolves $\frac{1.24}{9.25}$ gms. of solid,

and 100 gms. of water dissolve

$$100 \times \frac{1.24}{9.25} \text{ gms. of solid} = 13.4 \text{ gms. of solid.}$$

Solubility at 40° C. = 13.4.

EXERCISE 16.—Determine the solubility of washing-soda and alum at 20° C., 40° C., 60° C., 80° C.; also of nitre, chlorate of potash, and common salt.

Solubility Curves.—Since, as we have seen above, the same solid has so many different solubilities depending on the temperature, and as we may desire to know the solubility at any temperature, it is advisable to keep a permanent record of the solubility of any particular solid at *all* temperatures. To do this by experiments—which would have to be very numerous—would be extremely tedious, and so we adopt the following

graphic method. Experiment informs us that the solubility of potassium chlorate is as follows: At

10° C.	28° C.	50° C.	70° C.	90° C.
5.5	10.5	19	30	47.5

Obtain a piece of squared paper and proceed thus: Fix a point 0 in the position shown (Fig. 18), and mark off along the horizontal line points equidistant from one

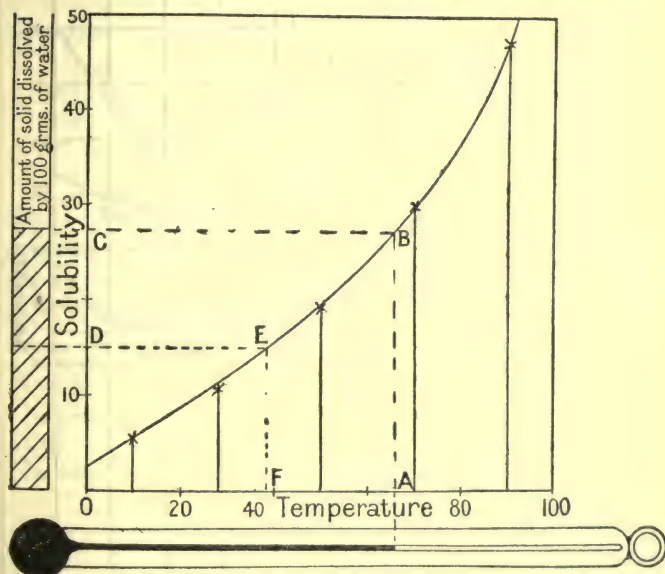


FIG. 18.

another, to represent thermometer readings, marking them 10°, 20°, and so on, to 100°. Similarly, mark off equal spaces along the vertical line, to represent the amount of solid dissolved by 100 grms. of water—i.e., solubilities—and mark them 10, 20, 30, etc. At the point marked 10° C. draw a vertical line 5.5 units long.

At the point 28°C . a vertical line 10.5 units; at the point 50°C . a line of 19 units, and so on for 70°C . and 90°C . Now draw a neat, even curve between these points—*i.e.*, not necessarily through all of them. All

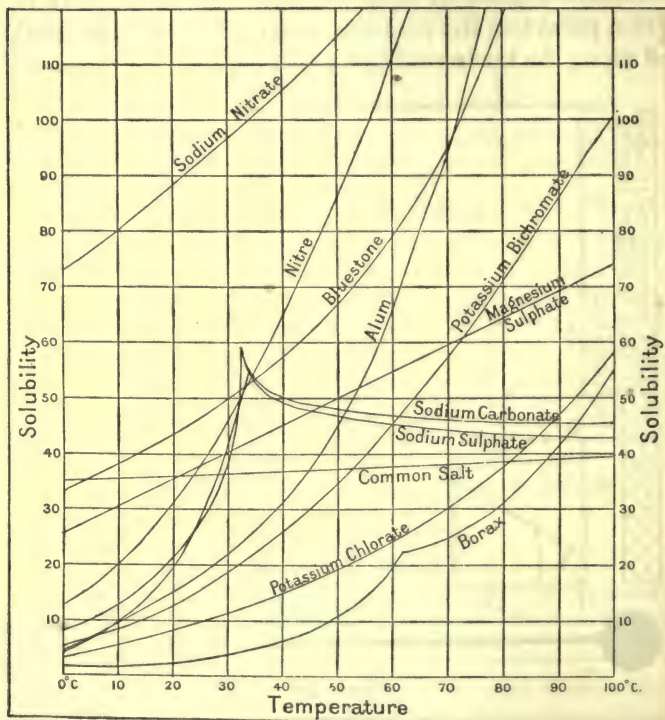


FIG. 19.

experiments possess some slight errors, so that some results may be too low and others too high. Such a curve is termed a **solubility curve**. Suppose, now, that we wish to know (1) the solubility of potassium chlorate

at 63° C. We look along the temperature-line to 63° C. (*A*), and ascend vertically until we reach the curve (*B*). The line *AB* measures the solubility and can be measured itself, or read off on the vertical scale = *OC*—i.e., 27. (2) The temperature at which the solubility is 15. Look along the vertical scale to 15 (*D*), and move horizontally to the curve (*E*). Then *DE* measures the temperature, which is seen to be on the scale = *OF*, or 39° C.

EXERCISE 17.—From your own results make a solubility curve for washing-soda and alum, and from the curve read off the solubility of each at 15° C., 30° C., 50° C., 75° C.

On p. 56 is printed a diagram of a few solubility curves, which shows clearly how the solubility rises with the temperature for most substances, and also how each substance is distinguished by its own solubility. Check your readings obtained from your own curves by readings taken from the curves given in Fig. 19.

CHAPTER IV

THE STUDY OF CRYSTALS

IN Expts. 27 to 29 we found that whenever a solid separated from a solution, it did so in a definite geometric shape—*i.e.*, as a crystal. (Why ?)

EXPT. 46.—To study the Action of Heat on Crystals.—Heat small quantities of the following crystals in small tubes, and observe closely all that occurs—bluestone, chlorate of potash, washing-soda, alum, Epsom salts, common salt, etc.—and verify the conclusion that **heat destroys the shape and colour of crystals**, and that in many cases steam escapes.

EXPT. 47.—To collect the Liquid from Crystals.—Half fill a 4-ounce retort with washing-soda (crystals), and distil (Fig. 12), stopping before the residue in the retort is quite dry. (Why ?) A colourless liquid collects in the flask. Taste it. What is it ?

EXPT. 48.—To examine the Liquid from Crystals.—**1. To find the Density of a Liquid.**—Weigh a clean dish, and by means of a pipette place in it 5 c.cms. of the liquid obtained in Expt. 47. Reweigh and calculate the weight of 1 c.cm.; it is 1 gm. Repeat this experiment with pure water; 1 c.cm. of this also weighs 1 gm. Repeat it with any other liquid, and show that this does not weigh 1 gm. **The weight of 1 c.cm. of a substance is termed its “density.”**

2. **To find the Boiling-Point of a Liquid.**—If you have a fairly large amount of the liquid, proceed to find its boiling-point as follows: Place some of this liquid in a boiling-tube, and fit the latter with a softened cork bored with two holes. Through one of these pass a thermometer, so that it is above the liquid, as shown, and through the other a right-angled tube. Heat the liquid and observe the thermometer. The temperature rises until the liquid boils, and then remains steady at 100°C . Repeat this experiment with pure water and other liquids, and show that only **pure water boils at 100°C** . If the liquid is inflammable, the apparatus shown in Fig. 13 is used.

It may chance that you have insufficient liquid to perform the above experiment. In such cases the **boiling-point of a small quantity** of a liquid can be carried out thus: Obtain a piece of thin-walled glass tubing, $\frac{1}{4}$ inch in diameter, and about 10 inches long. Seal one end, as in Expt. 24, and bend it, at a point 3 inches from the sealed end, into the U shape shown (Fig. 21).

Pour into the open end sufficient mercury to fill about 4 inches of the tube, and then enough of the liquid to cover the mercury surface to a depth of $\frac{1}{4}$ inch. By inclining the tube, transfer the liquid and mercury to the closed limb, allowing the air in it to escape. The tube will now appear as in the figure. N.B.—No air must be left in the closed limb. Attach a thermometer to the U-tube with india-rubber bands, and immerse the whole in a boiling-tube containing strong brine.

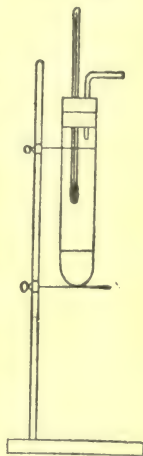


FIG. 20.

The whole of the closed limb must be immersed. Heat slowly, stir well, and observe the mercury level. When the level of the mercury is the same in both limbs, read the thermometer. Allow the whole to cool, and when the mercury is again level, again take the temperature. The average of these two temperatures is 100°C .—*i.e.*, the *boiling-point*. Verify this by repeating the experiment with pure water and other liquids whose boiling-points you know.

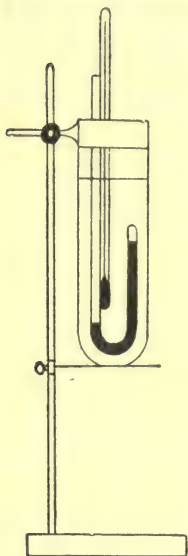


FIG. 21.

3. **To find the Freezing-Point of a Liquid.**—Into a test-tube place a little of the liquid obtained from the crystals, and a thermometer. Place this in a beaker containing a freezing-mixture of crushed ice and salt, and observe the stationary temperature; it is 0°C .—its **freezing-point**. Repeat this with pure water; this also freezes at 0°C ., whilst similar experiments with other liquids show that these freeze at very different temperatures. Thus it seems that many

solid crystals contain **water**, and that immediately this water is driven off they lose their shape and colour.

EXPT. 49.—To show that the Addition of Water often re-forms a Crystal.—Heat some blue copper-sulphate crystals in a dish until colourless and shapeless. When cold, add water, and observe the reappearance of the blue colour. Warm until dissolved and set aside to cool, and notice how the original shape is regained.

This water that is bound up in a crystal, and which gives it its colour and shape, is called the "water of crystallization"; but we must remember that all crystals do not contain water. Do all crystals that possess water of crystallization contain the same amount of this ?

EXPT. 50.—To find the Amount of Water of Crystallization in any Weight (say 100 gms.) of a Crystal.—Weigh a clean crucible and lid. (Why ?) Add about a gram of washing-soda and reweigh. Gently heat the crucible and contents on a sand-tray until apparently dry; cool and reweigh. Reheat, cool, and reweigh until of constant weight. (Why ?)

RESULTS—

Weight of crucible, lid, and crystals = 29.52 gms.

Constant wt. of crucible, lid, and dried solid = 29.14 „

∴ Weight of water = 0.38 gm.

Weight of crucible, lid, and crystals = 29.52 gms.

Weight of crucible and lid alone = 28.92 „

∴ Weight of crystals = 0.60 gm.

So that 0.60 gm. of washing-soda contains 0.38 gm. of water of crystallization;

∴ 1 gm. of washing-soda contains $\frac{0.38}{0.60}$ gm. of water of crystallization;

and 100 gms. of washing-soda contain

$$100 \times \frac{0.38}{0.60} \text{ gm. of water of crystallization.}$$

= 63.33 gms. of water of crystallization.

EXERCISE 18.—Repeat this experiment, using Epsom salts, copper sulphate, green vitriol, white vitriol, etc.

EXPT. 51.—To examine the Behaviour of Crystalline Substances on Exposure to the Air.—Weigh separate clean dishes containing clean crystals of washing-soda, calcium chloride, and magnesium chloride. Expose them to the air for some time, and observe them from time to time. The former crumbles to a dry powder (What has it lost?), and on reweighing is found to weigh less than before. Such substances that lose their water of crystallization when exposed to the air are called “efflorescent substances.” The calcium and magnesium chlorides, on the other hand, become covered with beads of moisture, and if left for a sufficiently long period, become quite liquid—*i.e.*, solutions—and on weighing, are found to increase in weight. (What is this due to?) Bodies like these which absorb water when exposed to the air are said to be “deliquescent,” and are very useful for drying gases, etc.

A third class of substances—*e.g.*, chlorate of potash, common salt, heated washing-soda, white copper sulphate, etc.—which do not contain water are termed “anhydrous.”

CHAPTER V

WATER

WE have already discovered that pure water is a great solvent, dissolving many different substances. All natural waters come, in the first place, from rain, and the question arises, Are natural waters pure ?

EXPT. 52.—To compare Natural Waters with Pure Water.—Evaporate 25 c.cms. of distilled, rain, spring, river, and sea water in separate weighed dishes with stirring-rods, and when the evaporation is completed, cool and reweigh. What do you discover ? Why does natural water contain dissolved solids ?

EXPT. 53.—To show that part of " Soil " is Soluble in Water.—Warm a small amount of soil with pure water, and filter off the undissolved portion. Evaporate the clear filtrate in a dish. What do you find ?

Is there anything else present in natural water ? Does water meet with any substances other than solids during its existence ?

EXPT. 54.—To show that Natural Water contains Dissolved Gases.—Fill a flask (*a*) (Fig. 22) with tap or spring water to the brim; fit it with a cork and bent tube, filled with water, as shown, by closing the lower end with the finger, and gradually working the cork into the neck of the flask. The end of the tube must only just pierce the cork. Place the whole on a retort-

stand and invert over the end of the delivery-tube in the trough a gas-jar or test-tube filled with water. Boil the water in *a* and observe what occurs. When no more gas collects in *c*, allow the apparatus to cool. What do you find? Test the gas in *c* with a lighted taper and lime-water. (What is it?)

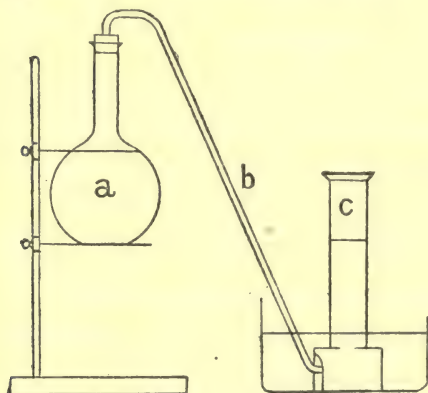


FIG. 22.

If this experiment is repeated with many natural spring - waters — *e.g.*, Perrier, Apollinaris, seltzer, etc. — much larger volumes of gas, consisting mainly of carbonic acid gas, are obtained. Thus, it seems that natural waters are impure, in that they contain solids and gases dissolved

in them. The water of the River Thames contains about 5.27 c.cms. of gas in every 100 c.cms. of water; the gas consists of 3.03 c.cms. of carbon dioxide, 1.50 c.cms. of nitrogen, and 0.74 c.cm. of oxygen; that of Bala Lake only 2.04 c.cms., consisting of 0.11 c.cm. of carbon dioxide, 0.64 c.cm. of oxygen, and 1.29 c.cms. of nitrogen.

Natural Waters.—The source of all water on the earth is, of course, rain, which is produced by evaporation from the surface of seas, lakes, rivers, etc., that is always proceeding, especially rapidly in the hot equatorial regions. This water vapour is cooled by rising and expanding or by flowing towards the cold

polar regions, and condenses to water or ice, falling as rain, hail, or snow. **Rain-water is thus naturally distilled water**, but before reaching the earth it has time to dissolve part of the air, dust, etc. The purity of rain-water will therefore depend upon the air through which it falls. In the country it is very pure, but in smoky towns it is far from pure. This rain-water either runs off the surface of the land, forming streams, or soaks into it, ultimately rising as a spring some distance away. (Which would be the purer, a stream or a spring?) In both cases it finally finds its way into rivers, lakes, and thence to the sea. During its contact with the earth (longer in the case of springs) it dissolves some of the solids over which it flows, the amount depending upon the nature of the rocks and the length of time it is in contact—*i.e.*, upon the rate and length of the river. If the rocks are hard, insoluble rocks, like granite, etc., very little is dissolved, and such is the case with the Scotch, Welsh, and North of England rivers. On the contrary, the rivers in the South of England flow over softer chalk or limestone, and contain much more solid in solution. Occasionally a river—*e.g.*, River Trent—contains a particular substance (in this case calcium sulphate), for it flows over a stratum of this substance. Note in the table on p. 66 the relation of the dissolved solids per 100,000 parts of water and the nature of the river basin.

Many other impurities are found in river-water from such sources as the refuse from towns on its banks, etc., of which it must be freed before it is fit for human consumption. Compare the following:

River Irwell, near Manchester, 7·8; at Manchester, 55·8.

River Mersey, at its source, 7·6; at Stockport, 39·5.

	Total Solids.	Length (Miles).	Nature of Basin, etc.
Rain (average) ..	2.9	—	—
Rivers and lakes (av.)	9.7	—	—
Springs (av.) ..	28.2	—	—
River Thames ..	28.0	215	Chalk, mainly; slow flow.
River Dee (Aberdeen)	5.7	90	Slate, sandstone, granite; rapid flow.
River Don (Aberdeen)	12.0	62	Ditto.
River Trent	70.0	180	56 parts of gypsum; a bed of this mineral lies in its course; slow flow.

Lakes manifest the same wide difference in composition. Some are practically the same as the rivers which supply and empty them, as—

Loch Katrine	28.3 parts.
Thirlmere	51.5 „
Lake Bala	27.9 „
River Dee	19 „ (10 to 28).

Others that have no outlet—*e.g.*, the salt lakes of Egypt, the Dead Sea, and the Elton Lake in Russia—act like enormous evaporating basins, accumulating the dissolved solids. Compare those given above with the following:

The Dead Sea ..	22,857 parts of solid.
The Elton Lake ..	27,143 „ „

Sea-water contains practically all the dissolved solids carried down by the rivers. This supply is more or less continuous, but evaporation takes place, removing the water and accumulating the solid—*i.e.*, the sea is

slowly becoming more "salty." In the larger oceans and seas the amount of solid dissolved is practically uniform, as will be seen from the table below:

Mid Atlantic	..	36	parts per 1,000 parts of water.
North Atlantic	..	37	" " "

Near the land it is less than farther out (Why ?), and in summer greater than in winter. (Why ?) In seas more or less land-locked the amount is high (Why ?), whilst in the polar regions the seas are less saline, due to the melting of the icebergs, which produces pure water, and the reduced evaporation. The equatorial seas are more salty, owing to the increased evaporation in those regions. Study the accompanying table:

Mediterranean Sea	..	40	parts per 1,000 parts of water.
British Channel	..	35	" " "
Irish Sea	..	34	" " "

Below is given a table showing the chief substances present in the British Channel per 1,000 parts of water:

Sodium chloride (common salt)	..	27.06
Magnesium chloride	3.67
Magnesium sulphate (Epsom salts)	..	2.30
Calcium sulphate	1.41
Potassium chloride	0.77
Magnesium bromide	0.03
Calcium carbonate	0.03

Total	35.27
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The above deals mainly with surface-water. We will now discuss the water which soaks into the earth and reappears as **springs**. Here again we find wide differences, varying from mere traces to amounts

greater than is found even in sea-water, according to the nature of the rocks through which it percolates and the length of time before its reappearance at the earth's surface. When a spring contains so much solid, etc., dissolved that it has acquired a taste, or is of value as a medicine, it is termed a **mineral spring**, sometimes a **medicinal spring**—*e.g.*, those of Harrogate, Droitwich, and Wiesbaden, etc., which contain salt; of Carlsbad and Marienbad, containing Epsom salts and sodium sulphate; Seltzer, Apollinaris, Perrier, etc., containing much carbonic acid gas; the **chalybeate** springs of Harrogate and Tunbridge Wells, containing iron; and the **sulphur** springs of Aix-la-Chapelle, Harrogate, etc., containing sulphides and sulphuretted hydrogen. Many springs coming from a considerable depth are much warmer than the air—*e.g.*, at Buxton, 28° C.; Bath, 47° C.; Carlsbad, 75° C.; and the geysers—in which the water is accompanied by steam-jets—in Iceland, New Zealand, and the Yellowstone Park, U.S.A. These are often termed **thermal springs**.

The Properties of Pure Water.—The water obtained in Expt. 30 by the distillation of impure water differs from all natural waters in that it does not contain solids or gases dissolved in it. It is **tasteless** (or flat), **odourless**, and practically **colourless**. We found in Expt. 48 that it has a **density of 1 gm. per c.cm.**, a **boiling-point of 100° C.** (when the pressure of the air is 760 mm.), and a **freezing-point of 0° C.** The latter characteristics we made use of to prove that the liquid obtained from crystals was water—*i.e.*, to identify it—and they are termed **physical constants**—*i.e.*, constants which attend certain physical changes—*e.g.*, of freezing and boiling. These constants are used for identifying,

as in Expt. 48, and as **tests for purity**, for only pure water has these. Verify this statement by finding the density, boiling-point, and freezing-point of impure water—*e.g.*, a solution of salt.

EXPT. 55.—**To study the Chemical Properties of Water.**—1. Add a small quantity of neutral litmus solution to some water, and observe that it does not become red, as it does with **acids** (Try one), nor blue, as it does with **alkalies** (Try sodium hydrate)—*i.e.*, water is a **neutral** liquid.

2. Add a little strong sulphuric acid to water in a test-tube, and observe how hot it becomes—a sign that **chemical union** has occurred.

3. Pour a small quantity of water on to some freshly prepared **lime (quicklime)**, and notice the clouds of water vapour given off. The mixture becomes hot—*i.e.*, chemical union has again occurred, **slaked lime**, or calcium hydrate [**A**] being formed.

4. We found in Expts. 27 and 28 that water unites with many colourless, shapeless bodies, and forms well-defined, coloured crystals. For this reason anhydrous (white) copper sulphate is often used to **detect** the presence of water in a liquid, etc.

Other properties will be discussed later, when the question, Is water simple or compound? has been settled. Heat, which, as we have seen, decomposes many compounds, merely causes water to boil away in the form of vapour or gas. Let us try the effect of another form of *energy* (not matter)—*viz.*, electricity.

EXPT. 56.—**To study the Effect of Electricity on Water.**—In the **voltameter** shown (Fig. 23), which consists of the upper half of a bottle closed by a cork,

through which pass two platinum wires carrying two platinum plates, pour some pure water. Connect the platinum wires to the terminals of an electric (galvanic) battery, and observe that little or no action seems to occur. Now add some dilute acid—*e.g.*, sulphuric acid. Immediately bubbles of gas are given off from each platinum plate. (What are these gases?) Fill two test-tubes with water and invert one over each platinum plate and collect a tube of each gas. Test the gas over *A* with a glowing wood ember; it is re-

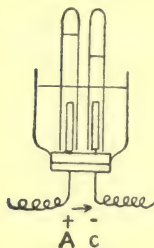


FIG. 23.

kindled—*i.e.*, the gas is **oxygen**. Test the gas over *c* by applying a light; the gas burns with a slight explosion—it is **hydrogen** [B]. Moreover, we observe that twice as much hydrogen collects as oxygen in volume. (Do these gases come from the water or from the acid?)

Test the liquid from time to time with blue litmus-paper. Even after a very long time the litmus is still reddened

—*i.e.*, the acid remains—and, although it may be necessary constantly to add water, no more acid is needed. Careful experiments show that when the action is over, as much acid remains as was added, so that **the gases must have come from the water, and not from the acid.**

Whenever, as in the above experiment, a **chemical decomposition is effected by electricity**, the process is termed **electrolysis**. The platinum plates, *A* and *c*, bearing or carrying the electricity into and out of the liquid, are called **electrodes**. The electrode *A*, where the current **enters** the liquid, is termed the **anode** (or +ve); and *c*, where the current **leaves**, the **cathode** (or -ve).

Thus we learn—

1. That electricity decomposes compounds.
2. That water is a compound of hydrogen and oxygen in the proportion of two volumes of the former to one volume of the latter.

All efforts to obtain any simpler substances from hydrogen and oxygen have so far failed; they are therefore regarded as **elements**, and we proceed to discuss them in detail in the following chapters.

CHAPTER VI

HYDROGEN

CHEMICAL symbol, H. Atomic weight, 1. Molecular weight, 2. Density = 1 (or 0.09 gm. per litre).

The name **hydrogen** means **water generator**, and its origin is clearly seen from the last experiment, and Expt. 59 (4) below, for when hydrogen burns, water alone is formed, as was first demonstrated by Cavendish.

EXPT. 57.—To study the Action of Metals on Dilute Acids.—In separate test-tubes, place fragments of zinc, iron, magnesium, etc., adding to each some dilute sulphuric acid. Warm gently if no visible action occurs, and apply a light to the liberated gas. Repeat, using various dilute acids—*viz.*, hydrochloric, nitric, acetic, and so on. Do you learn that **dilute acids, except nitric acid, react with many metals, liberating hydrogen?** [A]. Evaporate each of the liquids, after decanting or filtering, to a small bulk, and cool. Observe the crystals formed; these solids are known as **salts**.

Under special conditions, which will be discussed later, even nitric acid acts in a similar manner. We may therefore—

1. Define an acid as a substance containing hydrogen, capable of being replaced by a metal, to form a salt.

2. Use **dilute acids**, except nitric acid, and the **metals** zinc, iron, or magnesium, for the **preparation of hydrogen** in bulk.

Caution.—*The hydrogen apparatus must not be near a lighted gas-jet, nor must a light be brought near the end of the delivery-tube until the permission of the instructor has been obtained.*

EXPT. 58.—To prepare and collect Hydrogen from Acids.—Since an acid is to be added, a thistle funnel is required, which must always reach *below the liquid*. (Why?) Any suitable bottle or flask will do, a convenient form being a two-necked bottle (Woulffe's), as shown (Fig. 24). Zinc is generally employed, for it is less expensive than magnesium, and yields a purer gas than iron. Place in the Woulffe's bottle (*a*) a moderate quantity of granulated zinc; fit it with two softened corks.

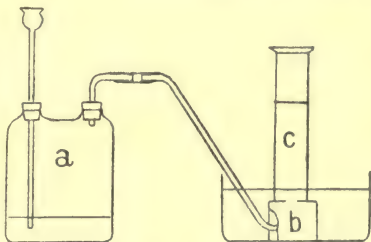


FIG. 24.

Through one pass a thistle funnel, and through the other a delivery-tube, bent as shown. The end of the delivery-tube dips under a beehive shelf (*b*), standing in a trough of water, and supporting a gas-jar (*c*) filled with water. Pour dilute sulphuric acid down the thistle funnel, and collect several jars of the gas. When a jar is full, cover the end with a greased, ground-glass plate, and remove it from the trough. Apply a light to the mouth of each jar, and notice that the first two or three *explode* loudly, but that afterwards the gas burns quietly. (Why?) Collect several jars of the gas and proceed to examine its properties as below.

EXPT. 59 (i.).—To examine the Properties of Hydrogen.—Examination informs us that hydrogen is colourless, odourless, and tasteless.

1. Since the gas bubbled through and was collected over water, it must be **insoluble in water**.

2. Add neutral litmus solution; it remains unaffected—*i.e.*, the gas is a **neutral gas**.

3. Plunge a lighted taper up into a jar of the gas; the gas burns, but the taper is extinguished, so that the gas does **not support combustion** (or burning), although it is a **combustible** substance—*i.e.*, one that burns itself. Observe the “dew” deposited [B] on the sides of the jar.

4. Place a dry jar on the top of a jar of hydrogen and remove the plate. Separate and test each jar. The upper jar now contains hydrogen—*i.e.*, **hydrogen is lighter than air**, bulk for bulk—*i.e.*, **is less dense**. Again, notice the “moisture” produced. These and other experiments show that **hydrogen burns, producing water**, and, as we learned in Expt. 56, must therefore have united with oxygen. (Where does this come from?)

The fact that hydrogen is less dense than air may be illustrated by several experiments.

5. The gas may be collected by turning the delivery-tube upwards and holding a jar of air over it. When you think the jar is full, remove it and test it with a light. (Does the gas ever burn without an explosion?) This is called “collecting” by **displacing air downwards**.

6. Balance two large beakers inverted, as shown (Fig. 25), and pour hydrogen upwards into A. Observe that B is now heavier than A—*i.e.*, air in B is heavier than hydrogen in A.

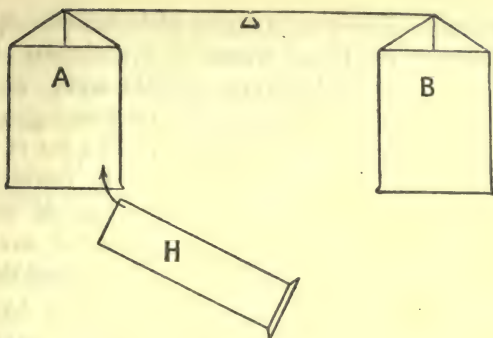


FIG. 25.

7. Owing to the small density of hydrogen (it is roughly one-fourteenth of the density of air), it is often used for filling balloons, etc. This may be illustrated as in 8 and 9.

8. Attach a small india-rubber balloon to the end of the delivery-tube of your hydrogen generator, and when fairly distended, pinch the neck of the balloon and remove it immediately from the delivery-tube. Tie the neck tightly with cotton and set it free; it rapidly rises to the ceiling.

9. Attach a clay-pipe to the delivery-tube and dip the end into some good soap solution. Bubbles filled with hydrogen are easily formed, which rise rapidly, and they can be set on fire by applying a light.

10. **To find the density of a Gas—e.g., of Hydrogen.**
—For this experiment it is necessary to prepare **pure, dry hydrogen**, which is carried out in the following way: In the bottle (a) (Fig. 26) is placed some magnesium ribbon, and dilute hydrochloric acid is added. The pure but wet gas is then dried by passing it through

a U-tube (*b*), or, better, a series of U-tubes, containing a substance “fond” of water—*e.g.*, calcium chloride, phosphorus pentoxide, lime, caustic soda, or strong sulphuric acid, etc. Before collecting any gas, a portion is allowed to escape. (Why?) If a jar of the gas

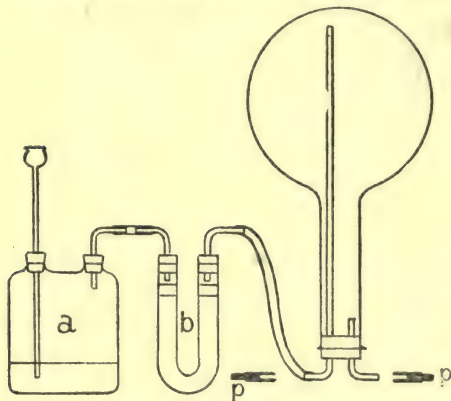


FIG. 26.

is required, the gas is collected over mercury. To find the density of hydrogen, use the apparatus shown (Fig. 26). This consists of a large flask, fitted with a cork, tubes, and plugs, as in the sketch. Fill the flask and tubes completely to the plugs (*p*)

with water, and pour this water into a graduated cylinder, thus finding its volume. Since 1 c.cm. of air weighs 0.001293 gm., we can from this calculate the weight of air in the flask, etc. Dry the apparatus and weigh the whole. By subtracting the weight of the contained air, find the weight of the flask, glass tubes, etc., as below, and retain these for future use.

Volume = 510 c.cms.

Now, 1 c.cm. of air weighs 0.001293 gm. (see Expt. 110).
510 c.cms. of air weigh $510 \times 0.001293 = 0.659$ gm.

Weight of flask, etc., and air = 252.240 gms.

Weight of air = 0.659 gm.

Weight of empty flask = 251.581 gms.

Attach the delivery-tube (Fig. 26) to *b*, and send a stream of pure dry hydrogen through the apparatus until full. (How will you know?) Plug up the tubes and reweigh. Calculate thus:

Weight of flask, etc., and hydrogen = 251.627 gms.

Weight of flask, etc., empty = 251.581 „

∴ Weight of hydrogen = 0.046 gm.

So that 510 c.cms. of hydrogen weigh 0.046 gm.

1 c.cm. of hydrogen weighs $\frac{0.046}{510}$ gm. = 0.00009 gm.

I.e., the density of hydrogen = 0.00009 gm. per c.cm.,
or 0.09 gm. per litre.

What sources of error are there in the above experiment? How would you suggest corrections for them?

Hydrogen is the least dense of all matter, and is therefore used as the **standard** for gas or vapour densities. Its **relative density** is therefore 1.

(ii.) Owing to its great “liking” for oxygen (it explodes when mixed with air or oxygen), hydrogen is capable of removing the former from many substances, such as the metallic oxides, etc. Such a process—viz., the removal of oxygen from a body—is termed **reduction**, and the agent effecting a reduction is termed a **reducing agent**.

To show that Hydrogen reduces Metallic Oxides to Metals.—Fit up the apparatus shown in the accompanying sketch (Fig. 27); it consists of a hard-glass tube (*a*) containing a porcelain boat (*b*) partly filled with copper oxide. Pass dry hydrogen through the apparatus by means of the tube (*c*) until all air is expelled

(Why? How will you know when this is so?); and then heat *b* strongly. (N.B.—When *thick* glass is

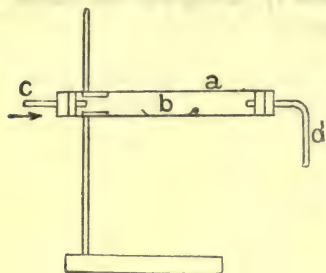


FIG. 27.

to be heated, it must be gently warmed at first, to avoid cracking.) Water drips from *d*, whilst the black oxide becomes converted into red metallic copper [C]. Thus, metallic oxides can be reduced to metals by hydrogen, which is itself oxidized to water—*i.e.*, hydrogen is a good reducing agent.

EXERCISE.—By weighing the boat (*b*) before and after the reduction and also when empty, find the weight of metal produced from 100 gms. of copper oxide. How will you know when the reduction is complete? Try other oxides.

RESULTS—

Weight of boat + oxide = 4.51 gms.

Weight of boat alone = 3.27 „

Weight of oxide = 1.24 gms.

Constant weight of boat + metal = 4.28 gms.

Weight of boat alone = 3.27 „

Weight of metal = 1.01 gms.

So that 1.24 gms. of oxide contain 1.01 gms. of metal.

1 gm. of oxide contains $\frac{1.01}{1.24}$ gms. of metal.

100 gms. of oxide contain $\frac{1.01}{1.24} \times 100$ gms. of metal,
= 80.64 gms. of metal.

I.e., the oxide consists of 80.64 per cent. of metal and 19.36 per cent. of oxygen.

EXPT. 60.—To examine the Residue left in the Hydrogen Apparatus.—After all action has subsided, filter the liquid and evaporate the clear filtrate to half its bulk and allow it to cool. Colourless needle-shaped crystals remain. (What are they ?) The zinc has disappeared, and the acid has been destroyed. (What suggestions can you make as to what has become of them ?)

EXPT. 61.—To discover what the “Residue” is.—Dissolve some of the crystals in pure water and place the solution in a beaker. Immerse two platinum electrodes, joined to the terminals of a battery, in the solution and pass an electric current for some time. Such an apparatus is called a *voltameter*. Examine the electrodes; the anode is unaffected, although bubbles of gas come from it, but the *cathode* is covered with crystals of **zinc**. Test them with dilute acid and the liquid with blue litmus, zinc, and washing-soda. The liquid is **acid**—viz., **sulphuric acid**—so that the substance formed in the hydrogen generator is a compound of zinc and sulphuric acid, and is therefore named **zinc sulphate [D]**. Whenever **sulphuric acid** is used, the **salt** formed is a **sulphate**; had we used **nitric acid**, a **nitrate** would have been formed; whilst **chlorides** result from **hydrochloric acid**.

QUANTITATIVE EXERCISES.

EXPT. 62.—To find the Volume of Hydrogen liberated by any Weight (say 1 Gm.) of a Metal.—Fit up the apparatus shown in Fig. 28. It consists of a small (2-ounce) flask (*b*), provided with a cork, through which passes a right-angled tube (*c*), pierced as shown. A delivery-tube is attached to *c* by means of india-rubber tubing, and a small tube (*a*) to *c* by a

piece of cotton. Weigh out 0.5 gm. of zinc and put it in the small tube (a); half fill b with dilute sulphuric acid and fit up as before. Test the apparatus, and

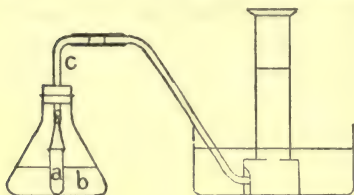


FIG. 28.

when it is air-tight, arrange to collect the hydrogen over water. Lower the tube (a) until the acid covers the zinc, and collect the gas given off. When all action is over, place a piece of gummed paper at the

water-level in the gas-jar; remove the latter, and fill up with water to the gummed mark. Pour the water into a graduated cylinder and measure its volume; calculate from your results as below:

RESULTS—

Weight of zinc = 0.61 gm.

Volume of gas = 210 c.cms.

So that 0.61 gm. of zinc liberates 210 c.cms. of hydrogen.

1 gm. of zinc liberates $\frac{210}{0.61}$ c.cms. of hydrogen

= 344.3 c.cms. of hydrogen.

Repeat this experiment with other dilute acids, and verify the statement that 1 gm. of a metal always liberates the **SAME VOLUME** of hydrogen, whatever the acid.

Again repeat the experiment with other metals, and prove that 1 gm. of **DIFFERENT METALS** liberates **DIFFERENT VOLUMES** of hydrogen.

EXPT. 63.—To find the Weight of Hydrogen liberated by any Weight (say 1 Gm.) of a Metal.—This is really the same as Expt. 62, since, if we know the volume of hydrogen and the density [as found in

Expt. 59 (10)], we can calculate the weight of hydrogen as follows:

From Expt. 62—

Weight of zinc = 0.61 gm.

Volume of hydrogen = 210 c.cms.

Since 1,000 c.cms. of hydrogen weigh 0.09 gm.,

1 c.cm. of hydrogen weighs $\frac{0.09}{1,000}$ gm.

210 c.cms. of hydrogen weigh $210 \times \frac{0.09}{1,000}$ gm.
= 0.0189 gm.

So that 0.61 gm. of zinc liberates 0.0189 gm. of hydrogen.

1 gm. of zinc liberates $\frac{0.0189}{0.61}$ gm. of hydrogen.

= 0.031 gm. of hydrogen.

By similarly calculating from your results of Expt. 62, verify the conclusions—

1. The same weight of a metal always liberates the same weight of hydrogen, whatever the acid used.

2. The same weight of different metals liberates different weights of hydrogen.

EXPT. 64.—To determine “directly” the Weight of Hydrogen set free by any Weight (say 100 Gms.) of Zinc.—Fit up the apparatus as shown. *b* is a 2-ounce flask containing dilute acid (1 c.cm. of strong sulphuric acid, or 8 c.cms. of strong hydrochloric acid, for every 2 gms. of zinc); *c* is a glass tube, pierced and bent as shown. Suspended in this by means of copper wire is a weighed zinc rod (*a*), about 5 gms.; *d* is a drying-tube, packed between cotton-wool with freshly heated, fine, granular, calcium chloride, to dry the escaping

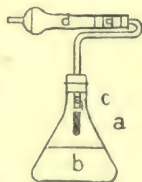


FIG. 29.

gas. The whole apparatus is now weighed, and the zinc is lowered, by rotating *c*, until its lower end is just immersed in the acid. The action commences, and hydrogen is **slowly** liberated, being dried ere it escapes through *d*. If the action is rapid, drying is incomplete, and the zinc must be gradually raised to slow the action down. When the action becomes too sluggish, the zinc is again lowered, and this is repeated until all the zinc is dissolved. The apparatus is again weighed, and the results calculated as below:

RESULTS—

Weight of zinc = 4.85 gms.

Weight of apparatus before = 120.58 gms.

Weight of apparatus after = 120.44 „

Loss, hydrogen = 0.14 gm.

So that 4.85 gms. of zinc liberate 0.14 gm. of hydrogen,

1 gm. of zinc liberates $\frac{0.14}{4.85}$ gm. of hydrogen,

and 100 gms. of zinc liberate

$\frac{0.14}{4.85} \times 100$ gms. of hydrogen = **2.89 gms. of hydrogen.**

[Accurate value, 3.09.]

Notes on Drying Agents.—1. The drying agent must be of good quality—calcium chloride of ordinary stock as a rule is inefficient: the greater proportion of its drying power has gone.

2. However efficient the drying agent, *time is the most important factor*. The above experiment to be satisfactory should take twenty-four hours, some $1\frac{1}{2}$ to 2 litres of the gas having to be dried.

We have already learned that water is a compound of hydrogen and oxygen, so that if we can only remove the latter, hydrogen should be obtained. This can be done, as in Expt. 56, by electrolysis, or by the action of certain metals, as below.

To study the Action of Sodium on Water.—(Caution.—*This metal must not come in contact with water on the hands, bench, etc.*).—Cut a small piece of the soft metal sodium and drop it on to the surface of water in a large trough. Observe that it melts, forming a globule that runs hissing about on the surface of the water until it has dissolved. Now place a larger piece of the metal in a piece of lead piping or copper gauze, drop it into the water and cover it immediately with a gas-jar of water. A gas collects; apply a light. The gas burns; it is **hydrogen**. (What has become of the sodium?) Test the liquid remaining (1) by rubbing on the hand—it feels soapy; (2) by adding a little red litmus to a portion—the litmus becomes blue—*i.e.*, an **alkali** has been formed; (3) by evaporating the remainder to dryness in a clean dish; a white solid remains. Leave a little of this exposed to the air; it soon becomes moist—*i.e.*, it is **deliquescent**. This white, caustic, deliquescent alkaline solid is **caustic soda (sodium hydrate, or sodium hydroxide)** [E]. The metal may also be dissolved in mercury, thus: Cut the metal into thin slices; place a slice in a little mercury in a mortar, and press it below the surface with the pestle (the hand should be covered with a dry duster) until it dissolves. Repeat this until all the sodium is dissolved. Now place this **sodium amalgam** in a crucible and stand it in water, collecting the gas as above.

The metal potassium acts similarly to sodium, only more vigorously, **caustic potash** [F] being formed.

EXPT. 65.—To make a Glass Jet.—Cut off a piece of suitable glass tubing about 6 inches long and rotate it near the top of the Bunsen flame, and finally in the

upper portion of the flame until it is fairly soft. Remove it from the flame and draw it out **steadily**. Its appearance should be as in Fig. 6. Draw the file smartly across the contracted portion and round off the ends as before.

EXPT. 66.—To study the Action of Steam on Magnesium.—Place some water in the flask (a) (Fig. 30)

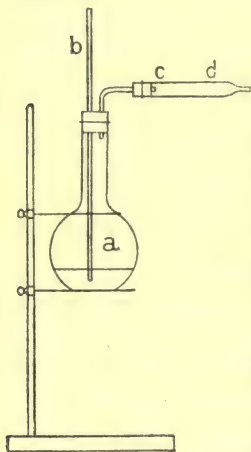


FIG. 30.

fitted with a safety-tube (b) and a right-angled delivery-tube (c). Attach to the latter a drawn-out test-tube (d) containing a little magnesium ribbon. Boil the water briskly, and when steam is issuing freely from d, dry, and then heat the magnesium strongly. It commences to burn brilliantly, forming the white solid **magnesia** (magnesium oxide), whilst a flame of burning **hydrogen** [G] is seen at the exit. If the gas is to be collected, the apparatus shown in Fig. 31 is employed. On testing, the gas proves to be **hydrogen**.

EXPT. 67.—To study the Action of Steam on Iron.—Pack an *iron* tube (Fig. 31, a) loosely with iron nails or filings (clean), and blow steam through it from a boiler. Heat the tube until it is red-hot, and examine the gas that collects; it proves to be hydrogen. Remove the delivery-tube from the trough, and when cool examine the residue in a. The residue is black and non-metallic in appearance, and resembles the iron scales of the blacksmith's-shop.

If Expt. 59 (ii.) is repeated, using this residue, iron is regained and water produced; it is indeed the black oxide of iron [H].

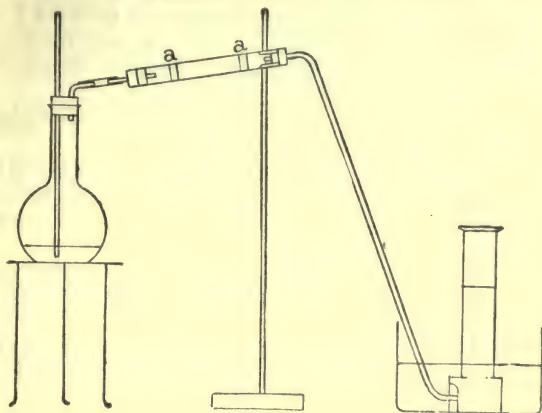


FIG. 31.

REVISION QUESTIONS.

CHAPTER I.

1. What do you understand by "matter" and "energy"? Give some examples of the effects of the latter on the former.
2. What do you know of the expansion of liquids? Mention any instances you can where use is made of such expansion.
3. If you required a thermometer to read very low temperatures, why would mercury be unsuitable? What liquid would you use? Give reasons for your answer.
4. Explain how you would verify the accuracy of the readings of a thermometer you had purchased.
5. Describe the construction of an ordinary mercury thermometer.
6. What do you understand by (1) heat, (2) temperature, (3) quantity of heat? Suggest a way of measuring the last mentioned.

CHAPTER II.

1. What evidence would you put forward in order to prove that a certain change was a chemical change? Which piece of evidence do you consider most decisive?
2. Define each of the following, and give an experiment illustrating each: Physical change, chemical change, solution, distillation, sublimation.
3. Suggest experiments to find out whether bluestone is a compound, element or mixture.
4. How could you separate the following mixtures: (1) Iron filings and salt; (2) iron filings and sand; (3) lead oxide and sulphur; (4) lead oxide and ammonium chloride; (5) iron filings and chalk.
5. If you were requested to examine the crystalline form of alum, and you were supplied with powdered alum, how would you proceed?
6. You are given some brimstone (or sulphur) and treacle. How would you get from it (1) some pure sugar; (2) some pure sulphur?

CHAPTER III.

1. Give experiments illustrating chemical and physical solution, and point out how you would distinguish them.
2. Describe how you would try to find a solvent for india-rubber. What is a solvent?
3. Explain clearly what you understand by "solubility" and describe experiments to show what it depends upon.
4. How would you ascertain the weight of nitre, of sulphur, and of charcoal in a pound of gunpowder?
5. Describe fully how you would determine the solubility of alum in water at 50°C .
6. You are supplied with a sample of mineral water. How would you find the amount of solid dissolved in a pint of such water?

CHAPTER IV.

1. Of what are most crystals made? Give experiments to illustrate your answer.
2. What do you understand by "water of crystallization"? Describe how you would determine the weight of water in 1 stone of alum.

3. Describe the various methods you know of obtaining crystals, and say what kind of changes occur in each experiment.
4. Explain what is meant by the terms "deliquescent," "efflorescent," and "anhydrous" substances. Describe one of each, and say what use is made of them, if any.
5. What liquids other than water have you used to obtain crystals? What kind of crystals are those you obtain? Describe the use of two of these solvents.
6. What do you understand by a "crystal"? Describe how you would prepare some *small* and one *large* crystal of bluestone. Which is the purer, and why?

CHAPTER V.

1. Write the biography of a raindrop.
2. Is natural water pure or impure? Give experiments to illustrate your answer.
3. What do you understand by "distillation"? Of what use is it? Describe how it is carried out.
4. What do you understand by "electrolysis"? Describe two experiments illustrating it, and say what you learn from them.
5. What do you understand by the "physical constants of water"? How could you determine them, and of what use are they?
6. Mention and illustrate by experiments the chief properties of pure water.

CHAPTER VI.

1. What does the name "hydrogen" mean? Give experiments which justify it.
2. What kind of hydrogen would you require in order to accurately find its density? Describe how you would prepare it and carry out the experiment.
3. What do you know of acids? What tests would you employ to find out whether a given substance was an acid or not?
4. Describe an experiment to ascertain the volume of hydrogen obtained when 10 lbs. of zinc is dissolved in dilute sulphuric acid. Mention all the precautions you would take in order to obtain an accurate result.

5. Describe an experiment by which you could obtain some lead from lead oxide. What other substances are produced and what property of hydrogen does this illustrate? Mention all precautions necessary.
6. Can hydrogen be obtained from substances other than dilute acids? If so, give experiments to show how this gas can be obtained from the substances you mention, and state what other products are obtained.

REVISION TESTS.

A.

1. Describe how you would make a mercury thermometer, and graduate it.
2. Distinguish between *suspended* and *dissolved* impurities in water. How can water be made fit for drinking (a) when it is muddy, (b) when it is brackish or salt.
[C.L.J., 1909.]
3. What do you understand by an "element" and a "compound"? What experiments would you perform on pieces of sulphur and marble in order to decide to which of these classes each belongs?
4. How would you prepare saturated solution of salt in water, and how would you ascertain the proportion of salt in a litre of the solution? Can one solid form more than one saturated solution with one liquid? If so, explain how this can be.
[C.W.B.J., 1906.]

B.

1. Describe experiments which illustrate the various effects of "heat" on "matter," and state the kind of changes that occur.
2. How would you find the amount of "rock"—insoluble in dilute hydrochloric acid—in a sample of mountain limestone?
3. What do you understand by a "solvent"? What liquids are usually employed, and how would you carry out experiments to find a solvent for a given substance?
4. Describe how you would proceed in order to find the amount of water of crystallization in a substance.

C.

1. Why is mercury usually employed in thermometers ? What other substances might be used ? Discuss their advantages and disadvantages.
2. Which is the purest form of natural water, and which is the most impure form that you know ? Describe how you could obtain pure water from the latter.
3. Name some common substances containing water of crystallization. How do crystals behave when exposed to the air ? Mention one example of each. Explain their behaviour, if you can.
4. Describe how you would prepare several jars of fairly pure hydrogen, and give experiments to illustrate its chief properties.

D.

1. If you required a *very accurate* thermometer, which would you choose, and why—(a) one with a large bulb or a small bulb ; (b) a fine bore or a large bore ; (c) a long stem or a short stem ?
2. Describe exactly how you would separate the components of (a) a mixture of salt and sugar ; (b) sand and sulphur ; (c) sulphuric acid and water.
3. What do you understand by the “solubility” of a solid ? Upon what does the solubility depend ? Give experiments to illustrate your answer.
4. How can you obtain water from hydrogen ? Describe how you would carry out this experiment, and all the precautions you would take. Can you obtain hydrogen from water ? If so, how would you proceed ?

E.

1. State all you know of the expansion of solids, and describe experiments to illustrate your answer. Mention any common instances you can in which such expansion is made use of or allowed for.
2. What kind of changes occur during solution ? Mention two instances of each, and describe how you would verify your answer.

3. What do you understand by a solubility curve? How would you make one for a substance—*e.g.*, alum? Of what use is it?
4. How would you determine *accurately* the volume of hydrogen given off when a weighed amount of zinc is dissolved in hydrochloric acid?

PRACTICAL TESTS.

1. Determine the amount of sand in the given mixture of sand and sulphur.
2. Examine the action of cold and warm sodium hydrate solution on the given metal aluminium. Collect any gas liberated, and identify it.
3. Determine the percentage of water of crystallization in the partially effloresced washing-soda supplied.
4. Find the amount of sugar and sand, in two ways, in the mixture supplied.
5. Find a solvent for "rouge." What kind of solvent do you find?
6. A mixture of sulphur and charcoal is supplied. Separate the components, and give them up for examination.
7. A mixture of copper and magnesium filings is supplied. Separate these as completely as you can, and leave samples of one, or both if you can, of the metals in a pure state.
8. Fit up a flask for generating hydrogen and collecting it by displacing air, and leave it for inspection.
9. The given substance is a mixture of potassium chloride and calcium sulphate. Obtain and show pure samples of each.
10. The substance supplied is a mixture of potassium sulphate and copper sulphate. Prepare pure specimens of each.
11. Find the volume of hydrogen obtainable from 1 Kgm. of the iron filings supplied when dissolved in dilute hydrochloric acid.
12. Find the weight of marble dissolved by 1 gm. of nitric acid.
13. Determine the weight of copper in 1 Kgm. of the cuprous oxide supplied.
14. Find the gain in weight when 100 grams of the magnesium supplied is heated in air.
15. Determine the solubility of lead nitrate in water at 10° C., 30° C., 50° C.

PART II

CHAPTER VII

OXYGEN

CHEMICAL symbol, O. Atomic weight, 16. Molecular weight, 32. Density, 16 (or 0.00144 gms. per c.cm.).

In the preceding sections we have learned that this element is present in air and water, and that it may be obtained from the latter by electrolysis (see Expt. 56). It can also be prepared from the air, as we shall discuss later. When an element unites with oxygen, the substance produced is termed an **oxide**—*e.g.*, in Expts. 66 and 67 magnesium oxide and iron oxide were formed.

EXPT. 68.—**To examine the Action of Heat on Metallic Oxides.**—Heat small quantities of as many oxides of metals as you can in small tubes, and test any gas liberated with a red-hot splinter. Few give off the gas that causes the wood to relight, for as a rule **metallic oxides do not give off oxygen when heated**. A few do—*viz.*, the oxides of mercury [A], silver, gold, and platinum. These lose oxygen and leave the metal, and they may therefore be used to prepare a small quantity of **pure oxygen**. The gas was discovered by Priestley in 1774 from mercuric oxide in this way.

Under certain conditions, oxides can be made which contain a larger amount of oxygen than usual; it is then termed a **dioxide** or a **peroxide**.

EXPT. 69.—To study the Action of Heat on Metallic Peroxides.—Repeat Expt. 68, using lead dioxide, manganese dioxide, barium peroxide, red lead, etc., and observe that all, more or less readily, give off oxygen. Examine the residue in each case; it is the oxide. So that **metallic peroxides, when heated, give off oxygen and form metallic oxides [B].**

How would you find out whether a given substance is the oxide or peroxide of a metal ?

Chemical names are so chosen that they indicate the elements present in a substance—thus:

Sodium hydr-oxide } denotes sodium, hydrogen, and
Sodium hydr-ate } oxygen.

Lead oxide denotes lead and oxygen.

Potassium chlor-ate denotes potassium, chlorine, and oxygen.

Zinc sulph-ide denotes zinc and sulphur.

Sodium sulph-ite denotes sodium, sulphur, and oxygen, and so on. So that whenever we see the termination **ate** or **ite** in a chemical name it indicates the presence of oxygen.

EXPT. 70.—To study the Action of Heat on Potassium Chlorate (Chlorate of Potash).—Heat a little of this white, crystalline substance *gently* in a tube, and observe that it melts, boils, and gives off a little oxygen. If the heating is increased, a more rapid evolution of gas occurs, and a white solid remains. (What is this ?)

EXPT. 71.—To examine the Residue obtained by heating Potassium Chlorate.—Dissolve a crystal of each of the following in a small quantity of pure water in separate test-tubes—viz.: (1) Potassium chlorate; (2) potassium chloride (Why ?); and (3) the residue

from Expt. 70. To each of these solutions add *one* spot of silver nitrate solution, and observe that a white precipitate occurs in Nos. (2) and (3) (What is this? What does it indicate? See Expt. 29); but that nothing seems to occur in the first case. Thus, **potassium chlorate**, when heated, liberates much **oxygen**, and forms **potassium chloride** [C]. This is by far the most convenient substance to use for preparing oxygen for experimental purposes.

EXPT. 72.—To prepare Oxygen.—As we have learned above, we may obtain oxygen by heating the **oxides of mercury, silver, etc., any metallic peroxide**, or from **potassium chlorate**; but still more easily may it be prepared as follows: Grind together in a *clean* mortar 1 part of manganese dioxide and 5 parts of potassium chlorate. This mixture is known as **oxygen mixture**. Heat a small portion of this in a small tube, and observe the *rapid* evolution of oxygen, even when gently heated. Half fill a hard-glass test-tube with the mixture, fit it with a softened cork and *wide* delivery-tube (Why?) as shown in Fig. 14, and *gently* heat the mixture. Oxygen is rapidly given off (if too rapidly, stop heating), and is collected over water. (What property does this indicate?) After several jars have been collected. examine its properties, as in Expt. 74 below. (Why does the gas come off more readily from this mixture than from potassium chlorate alone? What is left?)

EXPT. 73.—To examine the Residue from “Oxygen Mixture.”—Weigh out $\frac{1}{2}$ gm. of manganese dioxide, and add a fair quantity of potassium chlorate, in a porcelain dish. Heat until the oxygen is expelled; cool, and add water. Warm and separate as in

Expt. 43. When dry, weigh the insoluble; it is $\frac{1}{2}$ gm., and *black*, and further examination shows that it is manganese dioxide. Thus, in oxygen mixture the **manganese dioxide remains unchanged**, the potassium chlorate alone being decomposed at a much lower temperature.

EXPT. 74.—**To discover the Properties of Oxygen.**—Inspection of the gas informs us that it is **colourless, odourless, tasteless, and insoluble in water.**

N.B.—The gas prepared in Expt. 72 has a sweetish smell, for it is not so pure as that obtained in Expts. 68, 69, 70.

Caution.—*The non-metal phosphorus is so inflammable that it should be held in a pair of tongs, under water, and cut. It must never be touched with the fingers.*

1. Light a fragment of the **non-metal phosphorus** on a deflagrating spoon and place it in a jar of the gas. It burns brilliantly, clouds of white fumes being produced, the product being the **oxide of phosphorus (phosphorus pentoxide) [D]**. Add a spot or two of blue litmus solution; it is immediately reddened, showing that the oxide of phosphorus *dissolves* and forms an **acid (phosphoric acid) [E]**.

N.B.—*This solid is a strong poison, and great care must be exercised when using it.*

2. Light a fragment of **sulphur** and repeat the above experiment. The sulphur burns with a bright blue flame; a pungent-smelling gas, the **oxide of sulphur (sulphur dioxide) [F]**, is formed, which **dissolves** to form an **acid (sulphurous acid) [G]**.

3. Repeat this experiment, using a piece of charcoal (**carbon**). **Carbon dioxide [H]** is produced, which *dissolves*, forming the **acid, carbonic acid [I]**. The common name of this gas is carbonic acid gas.

Thus, it seems that **non-metals burn in oxygen to produce acidic oxides**—*i.e.*, oxides which, with water, form acids. Another name commonly used for these substances is **acid anhydrides**, which means literally **acids without water**, or acids from which water has been removed.

The name **oxygen** means **acid-producer**, and was given to the gas by Lavoisier.

Caution.—4. Repeat these experiments, using small quantities of the **metals** sodium, potassium, magnesium, iron, etc., without and with ignition before introduction into the gas. Sodium burns with a yellow flame; a white solid—**oxide of sodium**—is formed, which dissolves and turns red litmus blue—*i.e.*, is **basic** or **alkaline**. Potassium burns with a lilac flame, the white solid, **potassium oxide**, being formed, which is also **alkaline**. Magnesium burns with a dazzling white light, the white solid, **magnesia [J]**, or **oxide of magnesium**, being produced, which does not dissolve; whilst iron forms the insoluble black **oxide of iron [K]**. We shall find later that these are also **basic**, so that we may conclude that **metals burn in oxygen, forming basic oxides**. (Is the name “oxygen” a good one or not?)

These experiments also show that **oxygen supports combustion vigorously**, although it does not burn itself, is **incombustible**. The test we employed—*viz.*, the rekindling of a glowing wood splinter—shows this strikingly. **Moreover, oxygen is the only supporter of animal life.**

5. Test the gas with red and blue litmus; no effect is observed—*i.e.*, oxygen is a **neutral gas**.

EXPT. 75.—To prepare Pure Dry Oxygen.—Since oxygen is neutral, we may use any drying agent—*e.g.*,

calcium chloride, phosphorus pentoxide, caustic soda, or sulphuric acid—to dry it. Fit up the apparatus

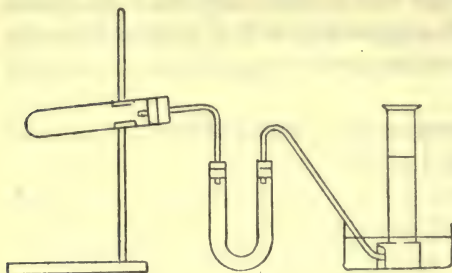


FIG. 32.

shown (Fig. 32), consisting of a hard-glass test-tube containing potassium chlorate or mercuric oxide, and attach to it a U-tube (or several U-tubes) containing calcium

chloride. Arrange to collect the gas over mercury, but before collecting, allow a portion of the gas to escape. (Why?)

EXPT. 76.—To find the Density of Oxygen.—Pass pure dry oxygen through the flask used in Expt. 59 (10) for some time; plug up the ends and weigh, calculating the density as before [Expt. 59 (10)], and verify the accuracy of the statement that the density of oxygen is 0.00144 gm. per c.cm. It will be thus seen that oxygen is $\frac{0.00144}{0.00009}$, or sixteen times as dense as hydrogen.

QUANTITATIVE EXERCISES.

EXPT. 77.—To find the Weight of Oxygen liberated by Heat from any Weight (say 100 Gms.) of a Substance.—When the substance is like mercuric oxide—i.e., leaves a liquid that boils away—this method cannot be employed. Weigh a crucible and lid; place in it about $\frac{1}{2}$ or 1 gm. of the substance and reweigh. Heat the whole on a pipe-clay triangle, as in Fig. 9,

at first gently, and finally strongly; cool and weigh. Reheat, cool, and reweigh until *the weight is constant* (Why?).

RESULTS—

Wt. of crucible, lid, and substance, before = 27.54 gms.
Constant wt. of crucible, lid, and substance,
 after = 27.32 „

∴ Loss = weight of oxygen = 0.22 gm.

Wt. of crucible, lid, and substance, before = 27.54 gms.
 Weight of crucible, lid, alone = 26.95 „

∴ Weight of substance = 0.59 gm.

Thus, 0.59 gm. of substance yields 0.22 gm. of oxygen;

∴ 1 gm. of substance yields $\frac{0.22}{0.59}$ gm. of oxygen,

and 100 gms. of substance yield $100 \times \frac{0.22}{0.59}$ gms. of oxygen
 = 37.29 gms. of oxygen, or 37.29 per cent.

EXPT. 78.—To find the Weight and Volume of Oxygen liberated from any Weight (say 100 Gms.) of a Substance

by Heat.—1. Weigh the hard-glass tube (a) (Fig. 33); place in it $\frac{1}{2}$ gm. of the substance and reweigh. Fit the tube with a softened cork and tube, bent as shown—i.e., the end passing up almost to the *top* of the gas-jar of water. (Why?) When *air-tight* (How will you know this is so?), heat until no

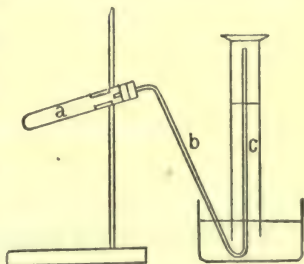


FIG. 33.

more gas is liberated. (How can you tell this?) Place a gummed strip of paper at the level of the water, after *allowing the whole to cool* (Why?), remove

the gas-jar and fill the part occupied by the gas with water, pour it into a graduated cylinder and note the volume. Disconnect and reweigh the tube and contents.

RESULTS—

Volume of the gas = 152 c.cms.

Weight of tube and substance = 12.54 gms.

Weight of tube alone = 11.92 „

∴ Weight of substance = 0.62 gm.

Thus, 0.62 gm. of substance liberates 152 c.cms. of gas;

∴ 1 gm. of substance liberates $\frac{152}{0.62}$ c.cms. of gas,

and 100 gms. of substance liberate

$$100 \times \frac{152}{0.62} \text{ c.cms. of gas} = 24,508 \text{ c.cms. of gas.}$$

Weight of tube and substance, before = 12.54 gms.

Constant wt. of tube and substance, after = 12.32 „

∴ Loss = weight of oxygen = 0.22 gm.

Thus, 0.62 gm. of substance liberates 0.22 gm. of gas;

∴ 1 gm. of substance liberates $\frac{0.22}{0.62}$ gm. of gas,

and 100 gms. of substance liberate $100 \times \frac{0.22}{0.62}$ gms. of gas

= 35.48 gms., or 35.48 per cent.

Since in this experiment we find the weight and volume of the same quantity of gas, we may use these results, to deduce the density of oxygen, thus:

152 c.cms. of oxygen weigh 0.22 gm.;

∴ 1 c.cm. of oxygen weighs $\frac{0.22}{152}$ „

= 0.00145 gm.

So that the density of oxygen = 0.00145 gm. per c.cm.

2. This experiment may also be conducted in the following manner: Attach the weighed glass tube containing the weighed substance to the aspirator (Fig. 34). Make this air-tight—when this is so, no water will run out—and proceed to heat. The end of the escape-tube *must always be under the water* in the jar (c). (Why?) When the decomposition is complete (How do you know?), cool, and calculate as above. Compare your results. (Which is the more accurate method? Why?)

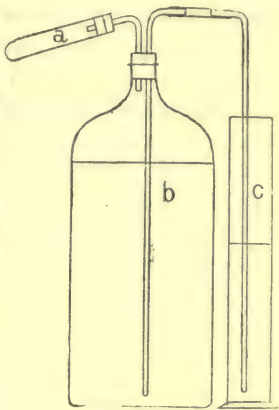


FIG. 34.

Since metallic oxides do not as a rule liberate oxygen when heated, we must proceed in quite another way.

EXPT. 79.—To find the Weight of Oxygen in 100 Gms. of a Metallic Oxide (Copper Oxide).—Weigh a clean crucible; place in it $\frac{1}{2}$ gm. or so of *clean* copper powder [B] and reweigh. Heat the whole strongly (Fig. 9) until the action is complete (How will you know?), cool and reweigh.

RESULTS—

Constant weight of crucible and oxide = 28.56 gms.

Weight of crucible alone = 27.46 „

∴ Weight of oxide = 1.10 gms.

Constant weight of crucible and oxide = 28.56 gms.

Weight of crucible and metal = 28.28 „

∴ Gain = weight of oxygen = 0.28 gm.

Thus, 1.10 gms. of oxide contain 0.28 gm. of oxygen,

1 gm. of oxide contains $\frac{0.28}{1.10}$ gm. of oxygen,

and 100 gms. of oxide contain $100 \times \frac{0.28}{1.10}$ gms. of oxygen
= 25.45 gms of oxygen, or 25.45 per cent.

EXPT. 80.—To find the Weight of Oxygen in 100 Gms. of an Oxide—e.g., Copper Oxide.—Weigh the hard-glass tube (*a*, Fig. 27) packed tightly with well-dried asbestos at *c*, together with a porcelain boat (*b*). Place in *b* about $\frac{1}{2}$ gm. of *clean* copper powder and reweigh the whole. Fit *a* with a cork and tube, and connect this to a supply of oxygen. Pass oxygen through the apparatus and gently heat the copper [**B**]. When all action is complete, cool and weigh. Reheat in a stream of oxygen, cool and reweigh until the *weight is constant*, and calculate as in Expt. 79. (Of what use is the glass-wool or asbestos in this experiment? Which is the better method, and why?) Try this with other metals—e.g., iron filings, zinc, magnesium, etc.

EXPT. 81.—To find the Weight of Oxygen in any Weight of a Metallic Oxide by Reduction.—Since most metallic oxides lose oxygen when heated in hydrogen, we may use Expt. 59 (ii.) to solve the problem. Repeat Expt. 59 (ii.), using copper oxide, and confirm the results of Expts. 79 and 80. Repeat, using iron oxide, etc., and confirm your results.

CHAPTER VIII

OXIDATION AND REDUCTION

IN the previous sections we have discovered (Expt. 74) that oxygen is an exceedingly active gas, and readily unites with hydrogen, metals, and non-metals. The product of the union of any substance with oxygen is termed an **oxide**, and the process is called **oxidation**.

Oxidation can be brought about in many ways, of which we have already studied the following—

1. By heating or burning the substance in air (Expt. 79).

2. By heating or burning the substance in the oxygen (Expts. 74 and 80). Experiment 80 may be used to prepare larger quantities of the oxides of the metals and, when suitably modified, of the non-metals.

To Oxidize Non-Metals.—

Place the non-metal in the boat (a) (Fig. 35) in the hard-glass tube (b), and pass a stream of dry oxygen through the apparatus. The oxide, if it is a volatile solid — *e.g.*, phosphorus pentoxide—is collected in a flask cooled by a freezing mixture, as shown. When the oxide is a gas it may be collected in dry gas-jars or as a solution in water in the flask.

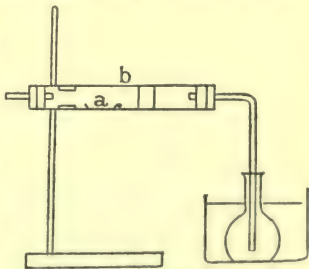


FIG. 35.

EXERCISE.—Prepare the oxides of zinc, iron, copper, carbon, sulphur, etc., in this way. The flask (c) will not be required (Why ?).

Other methods of oxidation will be met with subsequently.

The reverse of this process of oxidation—viz., **reduction—i.e., the removal of oxygen from a substance**—has already been mentioned. Thus, in Expts. 68, 69, and 70 the oxides of silver, mercury, gold, etc., were reduced to the metals themselves, the metallic peroxides to the metallic oxides, and potassium chlorate to potassium chloride, simply by *heating* them, oxygen gas escaping in all cases. In Expts. 59 and 81 we made use of the fact that hydrogen, by reason of its great “liking” for oxygen, could withdraw it from its union with substances of less affinity. Similarly, in Expts. 66 and 67, sodium, magnesium, and iron were used to remove oxygen from water—i.e., to reduce water to hydrogen. Such reductions were effected by using a **reducing agent—i.e., a substance very “fond” of oxygen**. Other substances, then, which unite with oxygen with vigour, besides hydrogen, should be reducing agents. Some of the commonest of these are coal (carbon), coal-gas, and carbon monoxide. It is obvious that in each case **the reducing agent is oxidized** by the oxygen it withdraws.

EXPT. 82.—To reduce Lead Oxide (Litharge) by Means of Carbon.—Mix together powdered charcoal or coal with about twice its bulk of litharge, and nearly fill a crucible with the mixture. Heat the whole strongly, stirring occasionally with an iron rod. Observe the tiny beads of metallic lead. After half an hour or so, pour the hot mass on to a clean iron tray; a button of lead is obtained. Now heat a little of the mixture

in a small tube, held in the mouth of a wider one, containing a little clear lime-water. The lime-water becomes "milky," or turbid, showing the presence of carbon dioxide. So that **carbon reduces lead oxide to lead**—*i.e.*, is a **reducing agent**—being **itself oxidized to carbon dioxide [B]**. Many metals—*e.g.*, iron, zinc, etc.—are manufactured in this way from their ores.

Coal-gas, which consists chiefly of hydrogen and compounds of carbon, is a very useful reducing agent.

EXPT. 83.—To reduce Copper Oxide by Coal-Gas.—Repeat Expt. 59 (ii.), but send coal-gas through the apparatus, *observing all the precautions taken there*.

EXERCISE.—Find the weight of copper and oxygen in 100 gms. of copper oxide, and deduce the amount of oxygen that unites with 1 gm. of copper. Also try lead oxide, iron oxide, tin oxide, etc.

In the above experiment water will be observed on the cool sides of the tube, and if the escaping gases are bubbled through lime-water, the presence of carbon dioxide will be indicated, showing that the coal-gas has been oxidized. What elements are present in coal-gas? What gases are present? Verify your answer by experiments.

EXPT. 84.—To reduce Lead Oxide by Means of the "Reducing" Blowpipe Flame.—Before commencing this experiment, **examine your Bunsen flame**. Close the holes at the base of the burner—the flame is large—with a blue inner cone and a yellow outer cone, giving out much light—*i.e.*, **luminous**. Now open these holes; again two cones are seen, but now practically no light is obtained—*i.e.*, the flame is **non-luminous**. (Why this difference? Does anything enter or escape at these holes?)

1. Hold a piece of smouldering touch-paper—made by soaking brown paper in a solution of nitre—near one hole, and close the other with your finger. Observe how the smoke is drawn *into* the Bunsen tube—*i.e.*, air enters when these holes are open. What does this air do?

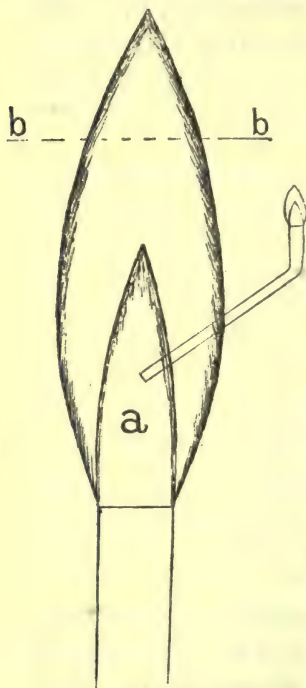


FIG. 36.

2. Hold a piece of glass tubing in the yellow portion of the flame; observe the **soot** deposited. Repeat this experiment with the non-luminous flame, and answer the above question after having burned a little soot in a jar of oxygen and tested the residual gas with lime-water.

3. Hold a short glass tubing in the inner cone, as at *a* (Fig. 36), and apply a light to the outer end. Observe how **unburned gas** can be drawn off here. Repeat this experiment in the outer cone.

4. Hold a match-head at *a*, and observe that the wood burns before the head bursts into flame. Try the same experiment at *b*. Do we not learn that (1) a **luminous flame** is made up of an inner cone

consisting of **cool, unburned gas**, and an outer, hot cone, containing **white-hot, unburned carbon**; (2) a **non-luminous flame** is made from this by the entrance of **air**, and consists of an inner cone of **cool, unburned gas** and a hot outer cone, which does not contain unburned carbon?

In the luminous flame of the Bunsen burner, then, near the top of the inner cone, we have (1) *unburned coal-gas*, (2) *hot, unburned carbon*—two very good reducing agents. This part of the flame is capable of effecting reduction, and is therefore called the **reducing flame**. In order to make it more convenient to work with, the flame is *gently* blown into the position shown (Fig. 37) with a blowpipe. No noise must be made. (Why?) At the point *R.F.*, any substance

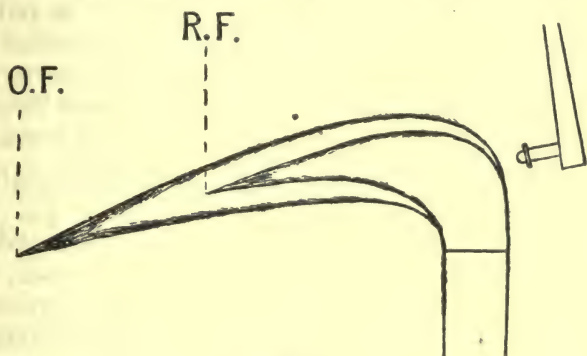


FIG. 37.

introduced is heated, whilst it is surrounded by hot, unburned coal-gas and carbon—*i.e.*, by reducing agents—and will lose its oxygen. *R.F.*, is therefore termed the **reducing flame**. At this point introduce some lead oxide on a piece of charcoal or porcelain, etc., and observe the ready reduction to lead. Now place this lead at the portion of the flame marked *O.F.*, where it is made hot whilst surrounded by air; here oxidation occurs and lead oxide is remade. This is the **oxidizing flame**.

EXERCISE.—Reduce tin and copper oxides in this way and then reoxidize the metal obtained.

CHAPTER IX

THE STUDY OF WATER—*Continued*

By the action of the metals sodium, potassium, magnesium, and iron, on pure water (Expts. 66 and 67) hydrogen was set free, which, since these metals are **elements**, must have come from the water. In the former cases **hydrates** of sodium and potassium were formed, whilst **oxides** of the latter remained. (How is this?)

EXPT. 85.—**To show that Caustic Soda contains Hydrogen.**—Warm a little water with a piece of aluminium or zinc foil; no gas is evolved. Add a piece of solid caustic soda to water and repeat the experiment. Apply a light to the gas; it explodes. So that sodium hydrate and aluminium or zinc liberate hydrogen; hence the name **sodium HYDR-ate** [A].

These experiments show that hydrogen can be expelled from water in **two** stages, and that, while sodium is only strong enough to expel **one** part, magnesium and iron are capable of expelling **both**. Since magnesium oxide only contains magnesium and oxygen, it follows that water contains **two or more "parts" of hydrogen united with oxygen**. This conclusion is borne out by the decomposition of water by electricity—*i.e.*, electrolysis (see Expt. 56) for two gases collected—*viz.*, **two volumes of hydrogen and one volume of oxygen**—which, if nothing else is present, is the composition of water by volume.

EXPT. 86.—To show that only Hydrogen and Oxygen are present in Water.—Fit up the apparatus shown for generating pure dry hydrogen, and attach a glass tube, bent as shown (a)(Fig. 38). When all air has been expelled from the apparatus (Why? How will you know when this is so?), light the escaping hydrogen and bring under it a dry jar of oxygen. Observe the “*moisture*” [B] deposited. Bring a dry jar of air under the jet and again observe the formation of “*moisture*.” The flame soon dies out, so that in order to be quite sure that this moisture is water, we must prepare a larger quantity for examination.

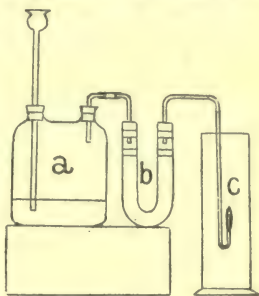


FIG. 38.

EXPT. 87.—To prepare Water by burning Hydrogen in Air [B].—Set up the apparatus as shown (Fig. 39), allowing the flame of the burning hydrogen to play upon the cold surface of a retort or flask, kept cold by a continuous flow of cold water.

A colourless liquid collects in the beaker. Examine this (1) with litmus: it is neutral; (2) by evaporation: no residue is

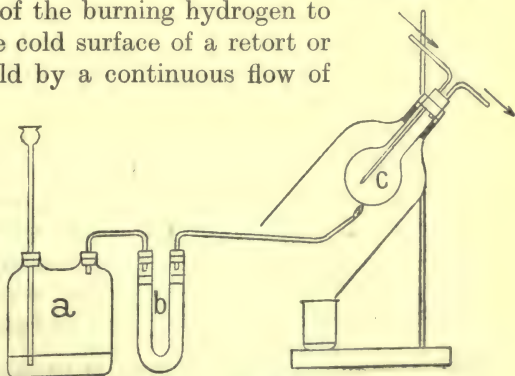


FIG. 39.

left; (3) by finding its boiling-point: it boils at $100^{\circ}\text{C}.$; (4) by finding its freezing-point: it freezes at $0^{\circ}\text{C}.$

(5) by finding its density: it is 1 gm. per c.cm.; (6) by its action on sodium: hydrogen is liberated; (7) by electrolysis: two volumes of hydrogen and one of oxygen are found. **It is, indeed, pure water.**

It was **Cavendish** (1731-1810), the true discoverer of hydrogen—although this gas had been known in a vague kind of way since **Paracelsus**—who first carried out the above experiments and indicated **Watt's** idea that oxygen and hydrogen united, producing *water, heat, and light only*. In 1781 he carried out a series of accurate eudiometer experiments (see Expt. 89), and proved that two volumes of hydrogen united with one volume of oxygen to produce water. This was confirmed by **Lavoisier** in 1786, who decomposed steam by iron, and found the composition of water by weight and by volume. In 1805 **Gay Lussac** and **Humboldt** showed that when two volumes of hydrogen united with one volume of oxygen, they formed two volumes of steam (Expt., p. 111); whilst **Dumas**, in 1842, carried out a series of accurate gravimetric experiments on the composition of water by weight.

Thus we are certain that water is made up of the two gases, hydrogen and oxygen, and of these only; and we may now use the results obtained by electrolysis and say that the **composition of water by volume is two volumes of hydrogen united with one volume of oxygen**. Since the densities of hydrogen and oxygen (see Expts. 59 (ii.) and 76) are respectively 0.00009 and 0.00144 gm. per c.cm., we may calculate the composition by weight—thus: 2 c.cms. of hydrogen weigh 2×0.00009 ; i.e., 0.00018 gm., which unites with 1 c.cm., or 0.00144 gm. of oxygen; or 1 gm. of hydrogen unites with $\frac{0.00144}{0.00018}$, i.e., 8 gms. of oxygen; so that the composition of water by weight is $\frac{1}{9}$ hydrogen and $\frac{8}{9}$ oxygen.

In Expts. 59, 81, etc., we reduced copper oxide by hydrogen, obtaining copper and water, and this may be used to determine, *by experiment*, the composition of water by weight, thus *verifying* the results obtained above.

EXPT. 88.—To determine the Composition of Water by Weight.—In *a* (Fig. 40), hydrogen is generated from zinc and dilute sulphuric acid, and passes through a solution of potassium permanganate in *b* to purify it, then through the U-tube (*c*) containing calcium chloride or other drying agent to dry it. This pure

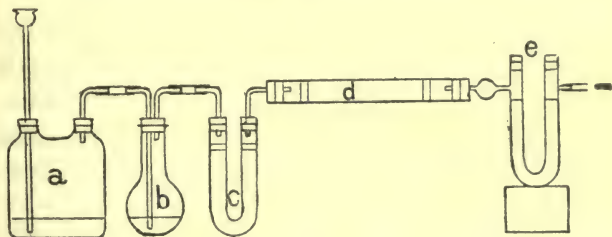


FIG. 40.

dry hydrogen passes through the weighed hard-glass tube (*d*), containing *pure, dry* copper oxide [*C*] (Why?), and thence through the weighed absorption-tube (*e*), containing calcium chloride. When all the air is expelled (Why?), the copper oxide is strongly heated. Reduction occurs and water collects in *e*. When the reduction is complete, *cool, while hydrogen is still passing* (Why?), and, when cold, disconnect *d* and *e* and reweigh separately; *d* has lost in weight, due to the oxygen removed; *e* has gained, due to the water formed.

RESULTS—

Weight of copper oxide and tube before = 45.62 gms.

Weight of copper oxide and tube after = 44.81 „

∴ Loss = oxygen = 0.81 gm.

Weight of tube *e* after = 17.65 gms.

Weight of tube *e* before = 16.74 „

∴ Gain = water = 0.91 gm.

Weight of hydrogen = weight of water – weight of oxygen = 0.91 – 0.81 = 0.10 gm.;

∴ of water, hydrogen forms $\frac{10}{91}$, i. e., $\frac{1}{9}$, by weight,

and oxygen forms $\frac{81}{91}$, i. e., $\frac{8}{9}$, by weight.

By electrolysis we found the composition of water by analysis. In the following experiment, as also in Expt. 88, we find this by synthesis.

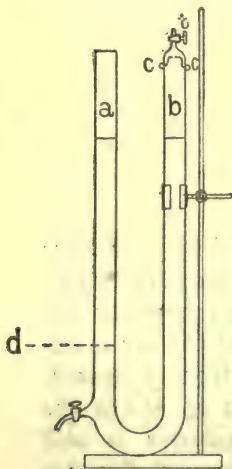


FIG. 41.

EXPT. 89.—To find the Composition of Water by Volume (Eudiometer).—We will now repeat Cavendish's experiment. The apparatus shown in the sketch (Fig. 41) is a U-tube eudiometer. It consists of a graduated limb (*b*), provided with two platinum wires (*cc*), and a tap (*t*). The limb (*a*) is not graduated, and also has a tap. The whole is filled with mercury and connected by the tap (*t*) to a generator of pure, but *not dry*, hydrogen. (Why?) By running out mercury from *a* and opening *t*, hydrogen enters. When about 20 c.cms. of the gas are introduced, *t* is closed, and the mercury in *a* is adjusted, so

that the levels in *a* and *b* are the same. (Why?) The

volume is now read. An oxygen generator is next attached to t and about 4 c.cms. of the gas introduced as before, the volume being read after the mercury has been again levelled. Mercury is now run out of T until its level stands low at d ; a is corked and a spark sent through the wire gap, cc . An explosion occurs; remove the cork and allow the apparatus to cool, adding mercury in a until the levels are once more the same. Note the volume and test the residual gas.

RESULTS—

Volume of hydrogen $= 21.5$ c.cms.

Volume of oxygen and hydrogen $= 26.4$ „

After explosion—

the volume of hydrogen left $= 11.7$ c.cms.

Whence,

volume of hydrogen used $= 21.5 - 11.7 = 9.8$ c.cms.

volume of oxygen used $= 26.4 - 21.5 = 4.9$ „

So that, **by volume**, 9.8 c.cms. of hydrogen unite with 4.9 c.cms. of oxygen; *i.e.*, 2 c.cms. of hydrogen unite with 1 c.cm. of oxygen.

****EXPT.—To find the Volumetric Composition of Steam.**—Since in the above experiment the temperature of the apparatus is considerably below 100° C., the steam produced rapidly condenses. It is obvious that to repeat Gay Lussac's experiment we must arrange that the temperature, etc., is such that the steam produced remains as steam or gas. This is effected in the following manner: The U-tube eudiometer used in Expt. 89 has the limb (b) containing the gases surrounded by a "jacket" (j) (Fig. 42), through which the vapour from a liquid boiling above 100° C. can be passed. Amyl alcohol boiling at 130° C. is generally

employed. The eudiometer divisions can be made more visible by red lead or other suitable means. The apparatus is first filled with mercury, and about 4 to 6 c.cms. of electrolytic gas—*i.e.*, the mixed gases

obtained by the electrolysis of water—is introduced. The jacket (*j*) is placed in position and the vapour sent through until the temperature is constant, the pressure is levelled, the volume read, and the temperature noted. The jacket is now removed and the gases allowed to cool, when the **explosion is carried out as in the previous experiment**. Replace the jacket and again pass the vapour until the thermometer again registers the same temperature as before. Again level the mercury and read the volume; it is found to be two-thirds of what it was before. Hence three volumes of mixed gases—*i.e.*, **two volumes of hydrogen and one volume of oxygen**—unite to pro-

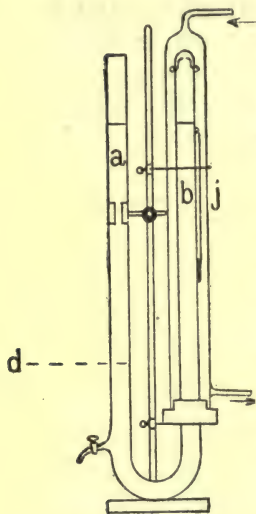


FIG. 42.

duce **two volumes of steam** at the same temperature and pressure. It is *essential* for the success of this experiment that the *mixed gases* and the *mercury* employed should be *dry*.

CHAPTER X

ELEMENTARY STUDY OF THE AIR

EXPT. 90.—To study the Action of Heat on Metals in and out of Air.—Place some clean fragments of the metals magnesium, zinc, lead, tin, iron, etc., in small tubes, and nearly fill the tube with asbestos, and heat. Cool and examine the residue. Repeat the experiment with the metals in crucibles or iron spoons, and see if you do not discover that **metals undergo a chemical change when heated in the air, but a physical change when air is absent.**

Some metals, like magnesium, burn, leaving an *ash*; others, like copper, *tarnish*; whilst others, like lead and tin, melt, and a *scum* or *dross* collects on the surface.

EXPT. 91.—To show that a Gain in Weight occurs when Metals are heated or burned in the Air.—In a clean crucible place some red copper powder [A] or turnings, and weigh. Heat this strongly on a pipe-clay triangle until black; cool and reweigh. (What do you learn?)

EXERCISES.—Repeat this with iron filings; magnesium, using a lid (Why?); tin, zinc, lead, etc., using a stirring-rod (Why?). Do your conclusions agree?

On examination, these tarnishes, etc., seem to be the same substances as those obtained by burning the metals in oxygen. Let us see if they are.

EXPT. 92.—To examine the Nature of Copper “Tarnish.”—Fit up the hard-glass tube (Fig. 27, *a*) containing a weighed porcelain boat, and place in it about $\frac{1}{2}$ gm. of pure copper powder and reweigh. Pass oxygen through the tube and heat strongly until black; cool and reweigh. Repeat until the weight is constant. Repeat this, but pass air through instead of oxygen, and prove that the gain in weight is the same in both cases.

EXPT. 93.—To show that Copper Tarnish contains the Metal and Oxygen.—Repeat Expt. 59 (ii.) with each of the residues in Expt. 91. Both are reduced by hydrogen, copper, and water being formed in each case—*i.e.*, copper tarnish is made up of copper and oxygen, or is **copper oxide**.

Similar results are obtained with the other metals, so that the **burning or tarnishing of metals in the air** is really a union of the metal with oxygen—*i.e.*, **oxidation**—and thus increasing in weight. It also proves to us that **air contains the active supporter of life and burning oxygen**.

EXPT. 94.—To study the “rusting” of Iron.—Clean a piece of sheet-iron thoroughly with emery-paper, and cut three pieces $\frac{1}{4}$ inch by 1 inch, and make the following experiments:

1. Boil some pure water and allow it to cool. In a clean test-tube place one piece of iron, fill up with this boiled water and cork it.

2. In a second tube place a lump of calcium chloride, or quicklime (For what purpose?) and a plug of asbestos. Cork it for a quarter of an hour, and then place in it a second piece of iron and cork again.

3. Take a third test-tube and place in it the remaining piece of iron and a few spots of water and cork. Leave

these three until the next lesson, and compare the results. Do not your experiments show that **iron rusts [C] in damp air only**? (How is this prevented in bridges, etc.?)

EXPT. 95.—To show that a Gain in Weight occurs during Rusting.—On a clean watch-glass place a few iron filings and weigh. Add a spot or two of water and set aside until rusted and *quite dry* (Why?) and reweigh. (What do you learn?)

EXPT. 96.—To examine Iron “Rust” [D].—Repeat Experiment 59 (ii.), using iron rust and reducing it in coal-gas; iron remains, and water is formed—*i.e.*, **iron rust is iron oxide**, so that **rusting is a case of oxidation**. (Why is it that certain metals never occur free in nature?)

EXPT. 97.—To study Burning in Air (of a Candle).—Place a piece of candle on a deflagrating spoon, and, after lighting it, place it in a dry gas-jar. Observe the “*moisture*” deposited. When the candle is extinguished, add a little clean lime-water and shake; the lime-water becomes turbid—a sign that **carbonic acid gas** is present. (Why does the candle go out?)

EXPT. 98.—To show that the Products of a Burning Candle weigh More than the Candle.—Fit the lamp-glass (*a*) (Fig. 43) with a cork carrying a small candle and pierced with holes. (Why?) Attach to its upper cork a U-tube containing soda-lime (a substance “fond” of both water and carbon dioxide), and weigh the whole. Attach an aspirator to this at *c*; light the candle and *quickly replace* in the lamp-glass. After about twenty minutes

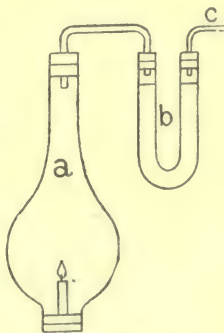


FIG. 43.

or so turn off the tap of the aspirator and observe the candle-flame. (What do you learn ?) Disconnect the aspirator and reweigh; a gain in weight is found.

Caution.—To study the Burning of Phosphorus.—Place a little of the yellow waxy solid phosphorus in a clean crucible on a plate. Ignite it with a hot wire and place over it a dry bell-jar. A *white solid* collects. Expose a little to the air; it is *deliquescent*. Add a little to water in a dish; it hisses, and the solution is acid to litmus. This solid was found in Expt. 74 to be the **oxide of phosphorus [E]**, or phosphorus pent oxide, and with water forms phosphoric acid.

Caution.—To show that a Gain in Weight occurs during the Burning of Phosphorus.—Place a little phosphorus

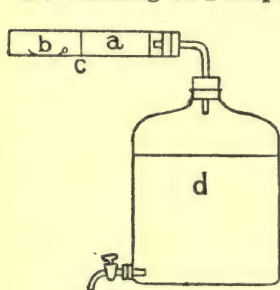


FIG. 44.

in *b* in the hard-glass tube (*c*) (Fig. 44), and pack the tube at *a* with asbestos and a little soda-lime. (Why ?) Weigh the whole and connect it to the aspirator (*d*). Ignite the phosphorus and turn on the tap. Observe that while the aspirator draws air over the phosphorus the latter burns, but that on turning off the tap it is immediately ex-

tinguished. (Why ?) When the phosphorus has burned, cool and reweigh; a gain in weight is found.

Similar experiments with other combustible substances show that—

1. Air is necessary for burning.
2. The products of burning weigh more than the original substance.
3. Oxides are formed—*i.e.*, burning is oxidation.

EXPT. 99.—To study Respiration (Breathing).—Fit up the apparatus shown (Fig. 45), placing in *a* and *b* fresh clean lime-water. Draw fresh air into the lungs; it enters through the lime-water in *a*, which remains clear. On expelling the air from the lungs, it escapes through the lime-water in *b*, which rapidly becomes milky, showing that carbon dioxide is formed in the body during breathing—*i.e.*, oxidation has occurred—and oxygen is consumed.

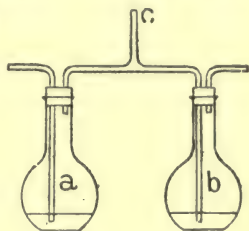


FIG. 45.

All these different processes—*viz.*, burning, rusting, tarnishing, and respiration—result in the formation of new bodies—oxides—by union with oxygen, which can only come from the air. (How much oxygen is there in the air? Is anything else present?)

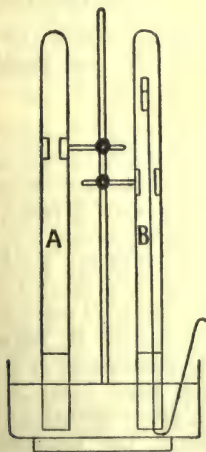


FIG. 46.

Caution.—To show that One-fifth of the Air is Active (Oxygen), and absorbed during Burning, etc.—Take two long glass tubes, *A* and *B*, and in the first place some wet iron filings or a bag of clean iron nails. In the other place a piece of phosphorus on the end of a long copper wire. Place them mouth downwards in beakers of water and examine them from day to day. As the iron-rust [*C*] and the phosphorus slowly

burns [*E*]*—i.e.*, unites with the oxygen—the water rises, until it fills one-fifth of the tube and then stops.

Remove the tubes, covering the ends with glass plates, and place in a lighted match; it is extinguished. So

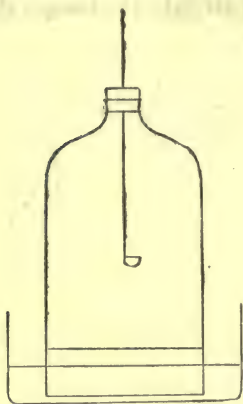


FIG. 47.

that by volume four-fifths of the air is inactive; it is the gas nitrogen, and one-fifth is oxygen, which is absorbed during burning, etc.

Caution.—To compare Slow and Rapid Burning.—Place a piece of phosphorus in a deflagrating spoon, passing through a cork fitted into a bell-jar of air, standing in water, as shown (Fig. 47). Light the phosphorus and place it in the jar, corking tightly, and compare the results obtained with those of the former experiment; they agree closely. So that **slow and rapid burning are the same chemically.**

Fire.—The phenomenon of "fire" has always possessed a fascination for the chemist. Aristotle regarded it as one of the fundamental elements, and various views have been brought forward from time to time as to its nature. **Sylvinus'** (1614-1672) suggestion that combustion and respiration were identical processes led (c. 1650) to the severance of the bonds between chemistry and medicine. After this date chemistry strove to become a self-supporting branch of natural science, and its history as a science begins properly with **R. Boyle** (1626-1691), who taught the "*acquisition of a knowledge of the composition of bodies—a study solely after truth along the lines of exact research.*" He too, however, often disregarded quantitative work, and although he knew the necessity of air for combustion, he believed it due to a fixation of a weightless element. We owe to him our knowledge of the elasticity and weight of the air, the air-pump, etc., which led him to formulate his law. He also knew the necessity of air for respiration, and in 1661, by publishing his book, the

"Sceptical Chemist," he destroyed all alchemical theories. He defined *elements*, showed that the union of elements gave compounds, and differentiated them from mixtures. Mayow (1668) showed that only a part of the air took part in combustion and respiration, and that the air that remained would not support combustion, and was somewhat lighter than common air.

Phlogiston Theory.—Stahl (1660-1734) suggested a "theory of combustion," which met with almost universal acceptance, known as the "*Phlogiston Theory*." According to his view, combustible bodies (and metals) contained a substance, **phlogiston**, which **escapes** during combustion. The amount of phlogiston varies according to combustibility, and can be transferred from one substance to another—*e.g.*, a calx (oxide of a metal) with phlogiston (carbon is rich in this element) produces a metal (rich in phlogiston). By means of this, Stahl grouped many reactions—*e.g.*, rusting, corrosion, respiration, and combustion—as identical processes, oxidation and reduction—by the removal or addition of phlogiston; but he *totally disregarded any quantitative results*. Black (1728-1799), on the contrary, concerned himself with quantitative work, and *urged the importance of the balance* in chemical work. About 1774 Priestley made his discovery of oxygen from oxide of mercury, and the astute mind of Lavoisier at once formed the correct view of combustion and allied processes. Although he made few discoveries, he swept away old prejudices and inaugurated a *period of quantitative work* by emphasizing the use of the balance as an instrument in chemical research. In 1772 he had shown that tin and phosphorus when heated in air absorb the air and gain in weight, whilst litharge (lead oxide) heated with carbon liberates a large quantity of a gas. Two years later, by heating tin in a sealed flask, he showed that the weight is constant until, on opening, air rushes in, and the weight increases. Learning from Priestley of the discovery of oxygen and of its properties, Lavoisier conducted his experiments on the heating of mercury in air (see below) in 1775, and identified the cause as the oxygen. This gave him the *true explanation of combustion*—*viz.*, *oxidation*—and overthrew the phlogiston theory. The fact that many substances (the non-metals) when they are burned in oxygen produce substances which, dissolved in water, form acids, led him to believe that oxygen was essential to the for-

mation of acids; hence the name. In 1777 he showed that phosphorus absorbs one-fifth of the air, and that the remaining four-fifths consist of an inactive gas—azote, or nitrogen. By his careful quantitative work he verified and gave a definite statement of the *law of conservation of matter*, and firmly established chemistry as an exact science.

Lavoisier's Experiments.—He placed some quick-silver or mercury in the flask (*a*) (Fig. 48), the end of which was drawn out and bent as shown. This passed up into a graduated bell-jar (*b*), containing air, and standing in a trough of mercury (*c*). The mercury in

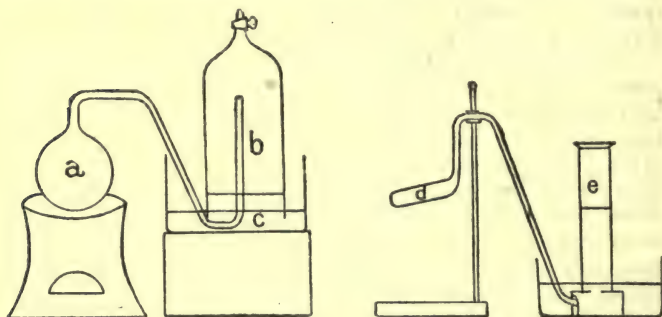


FIG. 48.

a was kept near its boiling-point for twelve days, during which time the air in *b* gradually diminished, until the mercury from *c* had risen and filled one-fifth of *b*. Meanwhile, a red scum [**F**] had collected on the mercury in *a*. This scum he collected and placed in the heating-tube (*d*), and strongly heated it. A gas was driven off and collected in *e*, whilst mercury [**G**] remained in *d*. From these experiments Lavoisier drew several important conclusions—viz.:

1. By replacing the gas obtained in *e* into the bell-

jar (*b*), air was remade: so that the gas obtained in *e* must have been obtained from the air.

2. On testing, the residue in *b* was found to be inactive **nitrogen**, whilst the gas in *e* was active **oxygen**.

3. When the mercury in *d* was added to that left in *a*, the original weight of mercury was regained.

Thus, Lavoisier learned that air is made up of two gases, oxygen and nitrogen, the former occupying one-fifth by volume, and is active, being absorbed during burning, etc., thus causing the formation of oxides, and the gain in weight observed; whilst the latter, which occupies four-fifths, is inactive, and not absorbed. Also, that during these changes no mercury or air was lost or destroyed. Thus, when a burning substance gains in weight, the air must lose a corresponding amount; or, if we consider both the air and the combustible substance, their total weight should be constant.

Caution.—To show that the Total Weight is Constant during Burning.—Fit the flask shown (Fig. 49) with a good cork, through which passes a tube and plug. Into the flask place a little asbestos or sand (Why?) and some phosphorus; cork tightly and weigh the whole. Now gently ignite the phosphorus, and when the burning is complete, cool and reweigh. (What do you learn?) Now, in a previous experiment we found that phosphorus increased in weight after being burned, so that the air loses exactly the same as the phosphorus gains, or the total weight remains constant.



FIG. 49.

If the tube is opened near the ear, air can be heard rushing in to fill the place of the oxygen absorbed, or if opened under water, this enters. So

that in all these processes oxygen is required, and it is the oxygen which unites with the substance, the increase in weight being the weight of the oxygen so used.

EXPT. 100.—To verify the Fact that “Matter is Indestructible.”—Heat a gram of copper in a porcelain boat in a stream of oxygen (see Expt. 80) until all metallic nature has [A] disappeared. Apparently the copper has been destroyed, only a black tarnish being left. Reduce this tarnish in hydrogen [B] or coal-gas, as in Expt. 59, until, on cooling and weighing, the weight is constant. Examine the residue; it is **copper**, and it weighs **1 gm.**, as before.

EXERCISE.—Repeat this experiment with lead, tin, etc., and compare your results.

EXPT. 101.—To show that a Candle behaves in the Same Way.—Place a small candle on the end of a wire and fix it in a cork in the flask (Fig. 49). Weigh the whole; remove the candle, light it, and *immediately* replace it in the flask, corking tightly. After a time the candle goes out, but on reweighing the weight remains the same. Here, indeed, the **candle is destroyed**, yet the **matter** of which the candle is made **is not destroyed**, but exists unchanged in the carbonic acid gas and water formed.

Now, **no experiment** is known which **contradicts** the facts observed by Lavoisier, and verified by the experiments above. This was recognized by Lavoisier as so important that he stated it as the **law of conservation of matter** in the following words:

1. **Matter cannot be created or destroyed ; or**
2. **The total quantity of matter taking part in any change remains unaltered when that change is completed.**

This important law is the very foundation of all the work of the chemist.

We have learned that air contains two gases, oxygen and nitrogen. A moment's thought will suggest *two others*, **carbonic acid gas** from fires and breathing, and **water**.

EXPT. 102.—To show the Presence of Carbon Dioxide in the Air.—Expose a dish of clean lime-water to the air, or blow a stream of air, by means of a pair of bellows, through lime-water. The lime-water becomes turbid, indicating that carbon dioxide is present.

EXPT. 103.—To show the Presence of Water Vapour in the Air.—1. Place on the bench a beaker containing ice or a freezing-mixture and examine the outside. In a short time moisture condenses, and even freezes, on the outside.

2. Place a piece of calcium chloride on a clean watch-glass and weigh. Expose this to the air of the room for half an hour; examine it and reweigh. Beads of moisture are observed, and the weight is found to increase.

There are many other substances present in the air, especially of towns, but in very much smaller amount, and are to be regarded more or less as accidental impurities.

EXPT. 104.—To prepare Oxygen from the Air.—This was first done by Lavoisier (see Expts., p. 120), for mercury, when heated in the air, absorbs oxygen, forming mercuric oxide. This oxide, when strongly heated, evolves oxygen gas, re-forming mercury. A substance which answers better than mercury is now used to *manufacture oxygen* by the *Brin's process*. Barium oxide, when heated moderately in air, absorbs

oxygen, forming barium peroxide [H]. This peroxide, like all other metallic peroxides, on being strongly heated, loses oxygen and re-forms barium oxide [I], which can be used repeatedly. It is essential that water and carbonic acid gas must first be removed from the air.

Nitrogen.

Chemical symbol, N. Atomic weight, 14. Molecular weight, 28. Density, 14 (or 0.00126 gm. per c.cm.).

This gas, as we have learned, occurs largely in the air. The name **nitrogen** means **nitre-producer**, from the fact that it occurs in nitre or potassium nitrate.

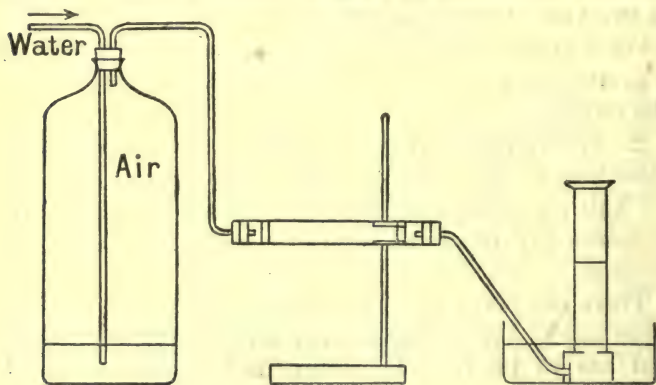


FIG. 50.

Owing to its inactive nature, whilst its companion, oxygen, is very active, it may be easily prepared from air by removing the oxygen with any suitable substance—*e.g.*, phosphorus, copper, etc.

Caution.—To prepare Nitrogen from the Air.—Small quantities were prepared on pp. 117-120; larger amounts may be prepared as below:

Fit up the apparatus shown (Fig. 50), consisting of an aspirator of air, into which water can be forced, thus sending a *slow* stream of air through the hard-glass tube, containing a boat holding phosphorus, then asbestos and soda-lime. (Why?) The issuing gas is collected over water, and may be examined as below. Other substances may be used instead of phosphorus [L], such as copper [A], iron, etc., although the action is much slower. The gas can be obtained in a purer condition from the white crystalline solid, ammonium nitrite, or, more conveniently, a mixture of potassium nitrite and ammonium chloride.

EXPT. 105.—To study the Action of Heat on Ammonium Nitrite.—Place a few crystals of this deliquescent solid in a small heating-tube and **gently warm**. Observe the clouds of vapour [J]. Collect a test-tube of the gas over water and test it (1) with a lighted taper, (2) with lime-water. (What is it?)

EXPT. 106.—To prepare and collect Nitrogen.—Place about 30 gms. of ammonium nitrite (or the mixture above mentioned) into the flask *a* (Fig. 51), and *gently* boil. A gas collects over water in the gas-jar, *b*. Proceed to examine it as below [J].

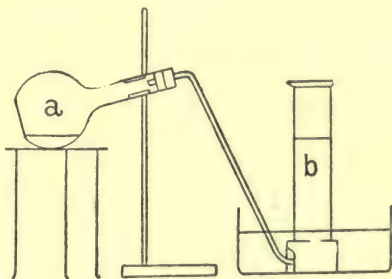


FIG. 51.

EXPT. 107.—To study the Properties of Nitrogen.—The gas, on examination, is observed to be **colourless**, **odourless**, and **tasteless**. Since it is collected over water, it is **insoluble** in water.

1. Test a jar of the gas with litmus; no effect is observed: it is a **neutral** gas.

2. Plunge into jars of the gas a burning candle, lighted phosphorus, sulphur, etc. They are all extinguished—*i.e.*, the gas is a **non-supporter of combustion**—neither does the gas burn—*i.e.*, it is an **incombustible**.

3. Shake up a jar of the gas with lime-water; no effect is observed.

4. An animal placed in the gas soon dies—*i.e.*, it is a **non-supporter of life**. It cannot be poisonous. (Why ?)

5. **To find the Density of Nitrogen.**—Pass the gas prepared as in Expt. 106 through a **U-tube** surrounded by cold water, and then through a **U-tube** containing calcium chloride (Why ?), and into the flask used in Expts. 59 and 76, and proceed, as before, to find the density. It is found to be **0.00126 gms. per c.cm.** Since hydrogen has a density of 0.00009 gm. per c.cm., nitrogen is $\frac{0.00126}{0.00009}$, or fourteen times as dense as hydrogen.

CHAPTER XI

THE STUDY OF THE AIR—*Continued*

The Composition of the Air, by Volume.—In so far as the nitrogen and oxygen are concerned, this has been roughly found in Expts., pp. 117, 118, and by Lavoisier's experiment, to be one-fifth oxygen and four-fifths nitrogen.

EXPT. 108.—To observe the Action between Air and an Alkaline Solution of Pyrogallic Acid (Pyro).—

1. Dissolve a little "pyro" in water and expose it to the air.

2. Dissolve a little "pyro" in potassium hydrate solution, and compare with the result in (1). To what is this due?

3. Take a dry test-tube and a tightly fitting cork. Put a gummed strip of paper at the lower edge of the cork and fill the tube with water to the mark. Pour this water into a graduated cylinder and note the volume. Pour into the tube 2 c.cm. of strong solution of caustic potash and then 2 c.cm. of strong "pyro" solution and cork immediately. Note anything that occurs. After five or ten minutes' shaking, invert in a trough of water and remove the cork. Measure the volume of liquid in the tube and find the amount of air absorbed. Test the residual gas. (What is it?)

EXPT. 109.—To find the Volume of Oxygen and Nitrogen in the Air.—*b* (Fig. 52) is a graduated gas-tube provided with a tap, and connected with a reservoir-tube

(a), both filled with water. By opening the tap and lowering *a*, air is drawn into *b*; the tap is turned off and *a* raised until the water-levels in *b* and *a* are the

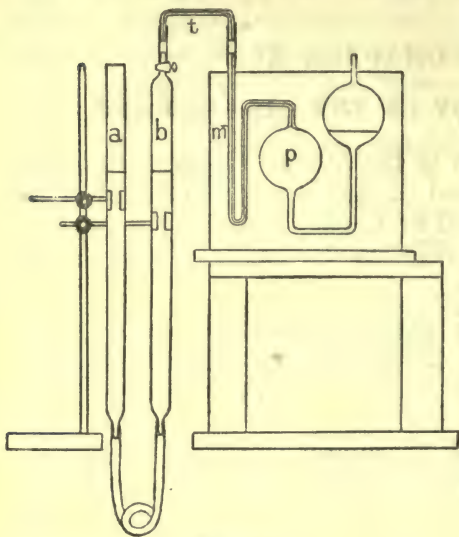


FIG. 52.

same and the volume of air read. It is now connected to the pipette (*p*) filled with alkaline pyrogallol solution to the mark (*m*). By opening the tap and raising *a*, the gas in *b* is forced into *p*, where the oxygen is dissolved, more rapidly if *p* is shaken after the tap is closed.

By lowering *a* and opening the

tap, the residual gas is drawn back into *b*, until the solution in *p* again stands at *m*. The tap is closed and the gas is then measured when the levels are once more the same. These processes of absorption and measuring are repeated until the volume of the residual gas is constant. (What does this show?)

RESULTS—

Volume of air = 21.2 c.cms.

Constant volume of residual nitrogen = 17.0 c.cms.

Volume of oxygen absorbed = $21.2 - 17.0 = 4.2$ c.cms.

i.e., oxygen forms $\frac{4.2}{21.2}$, or $\frac{1}{5}$ (approx.) by volume of air,
and nitrogen $\frac{4}{5}$. (What else is absorbed in the pipette?)

The accurate results obtained by many experiments point to **21 per cent. of oxygen and 79 per cent. of nitrogen by volume in dry air.**

Utilizing these results and the densities of oxygen and nitrogen, we can calculate the composition of air by weight; thus:

21 c.cms. of oxygen weigh $21 \times 0.00144 \text{ gm.} = 0.03024 \text{ gm.}$

79 c.cms. of nitrogen wgh. $79 \times 0.00126 \text{ gm.} = 0.09954 \text{ ,,}$

$\therefore \text{air} = 0.12978 \text{ gm.}$

So that 0.12978 gm. of air contain 0.03024 gm. of oxygen;

$\therefore 1 \text{ gm. of air contains } \frac{0.03024}{0.12978} \text{ gm. of oxygen,}$

and 100 gms. of air contain

$$\frac{0.03024}{0.12978} \times 100 = 23 \text{ gms. of oxygen.}$$

So that **by weight air consists of 23 per cent. of oxygen and 77 per cent. of nitrogen.** Indirectly, this also tells us the density of air approximately, for in the above 100 c.cms. of air weighs 0.12978 gm., or the density of air is 0.00129 gm. per c.cm.

EXPT. 110.—To find the Density of Air.—Repeat Expt. 59 (10), weighing the flask, exhausted by means of an air-pump, and then, when full of air, calculate its density. If accurately performed, it is found to be **0.001293 gm. per c.cm.,** or its density relative to hydrogen is $\frac{0.001293}{0.00009}$, or **14.4.**

The results *deduced* from Expt. 109 may be verified directly by experiment.

EXPT. 111.—To find, by Experiment, the Composition of the Air by Weight.—Fit up the apparatus shown in sketch (Fig. 53). *e* is a U-tube of calcium chloride, *d* one containing caustic soda, and *c* contain-

ing pumice and sulphuric acid. *b* is a long hard-glass tube packed with copper foil and weighed. *a* is a flask provided with taps (*t*), exhausted and weighed. *b* is made red-hot, and by turning on the taps a *slow* current of air is drawn through the tubes, *e*, *d*, *c*, where it has removed from it the water vapour and carbon dioxide, and then over the

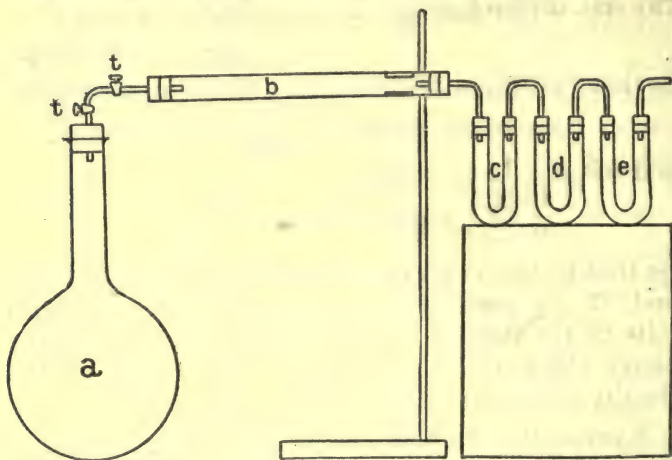


FIG. 53.

hot copper. This unites with and removes the oxygen, forming copper oxide, and gaining in weight, whilst the residual nitrogen collects in *a*. When *a* is full (How will you know?), the apparatus is allowed to cool, the taps are turned off, and *a* and *b* are re-weighed.

RESULTS—

Weight of *a* after = 47.62 gms.

Weight of *a* before = 47.25 „

∴ Gain = nitrogen = 0.37 gm.

Weight of *b* after = 37.61 gms.

Weight of *b* before = 37.50 „

∴ Gain = oxygen = 0.11 gm.

∴ Weight of air = $0.37 + 0.11 = 0.48$ gm., and percentage of oxygen = $\frac{0.11}{0.48} \times 100 = 23$ per cent. (approx.).

The average of many accurate experiments gives the results 23 per cent. of oxygen and 77 per cent. of nitrogen, by weight.

The amounts of carbon dioxide and water vapour in the air are so small compared with the quantity of oxygen and nitrogen that their accurate determination can only be made by using a much larger volume of air.

EXPT. 112.—To determine the Amount of Carbon Dioxide and Water in the Air.—Fit up the apparatus shown

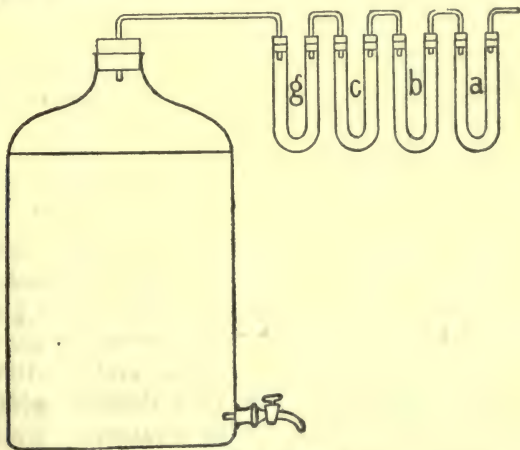


FIG. 54.

(Fig. 54), consisting of a large aspirator (10 or 20 litres) filled with water, attached to which is a U-tube (*g*),

containing pumice and strong sulphuric acid, called a guard-tube. (Why ?) This is joined to a U-tube (*c*), containing pumice and strong sulphuric acid, *b* containing caustic soda, and *a* pumice and strong sulphuric acid. *a*, *b*, and *c* are weighed before being connected with *g*. Water is slowly run out of the aspirator, its volume being noted, and this draws air through *a*, where the moisture is absorbed, *b* and *c*, where the carbon dioxide is retained. When a sufficient quantity of air has been drawn through the apparatus, *a*, *b*, and *c* are disconnected and reweighed, and the volume of water—*i.e.*, of the air—noted. The gain in weight of *a* gives the weight of water, and the gain in weight of *b* and *c* gives the weight of carbon dioxide in the volume of air employed.

Accurate determinations give the average results at the following values by volume:

Oxygen	20.66
Nitrogen	77.90*
Carbon dioxide	0.03
Water vapour	1.40
Total	100.00

Is Air a Mixture or a Compound?—Although the composition of the air is remarkably constant, due to the causes mentioned below, and to the general mixing up of the different parts of the air by diffusion and winds, yet from time to time, and in different places, the composition does vary slightly, while no compound exhibits any variation whatever. Furthermore, when we mix four volumes of nitrogen and one volume of oxygen, as did Lavoisier, the product behaves in every way like ordinary air, but no heat or

* Including the inert gases, argon, etc.

light is produced, such as we usually find when a compound is produced. When we dissolve a compound in water and evaporate the water away, we obtain the same substance again; but if we bubble air through water, boil out the dissolved gas and examine it, we find that instead of air containing 21 per cent. by volume of oxygen, we get a gas containing **33 per cent. by volume of oxygen**. Again, if air is subjected to great cold and pressure, it condenses to a blue liquid, which, if allowed to evaporate, gives us, not air, but, first of all, **pure nitrogen**, and afterwards **pure oxygen**. This is quite different to the behaviour of a compound like water, for example. For these and many other reasons we believe that air is merely a **mixture** of gases, and not a compound.

Life-Processes and the Composition of the Air.—We have already discovered that oxygen is the only supporter of animal life, and that carbon dioxide is produced during respiration. We now proceed to examine these processes in more detail.

EXPT. 113.—To determine the Composition of Respired Air.—Pass 50 c.cms. of expired air into the gas-tube (*b*) of the apparatus used in Expt. 109, collecting the last portion from the lungs. After measuring its volume, attach it to a pipette containing caustic soda, and determine the absorption as before. (What does this give?) Disconnect this pipette and determine the oxygen and nitrogen as in that experiment. Tabulate your results and compare them with the results obtained for pure air. Thus, **animal respiration consists in continuously removing oxygen and supplying carbon dioxide to the air.** (Why are animals warm?)

EXPT. 114.—To study the Behaviour of Plants.—You have probably noticed in an aquarium or pond bubbles of gas clinging to leaves, etc., of various plants growing under water. If not, attach a common

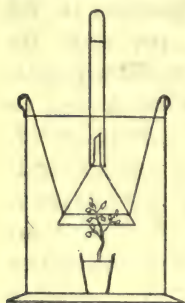


FIG. 55.

water-plant—*e.g.*, a pond weed, etc.—to a stone and plunge it under water in a deep vessel and expose it to the sun. If the water has carbon dioxide bubbled through it for a short time, the experiment works rather more rapidly. In a short time bubbles of gas come off, which may be collected as shown (Fig. 55). Test this gas with a glowing piece of wood; it relights, for it is mainly oxygen.

Thus, green plants absorb and remove carbon dioxide from the air, and return pure oxygen for fresh animal consumption.

We have seen above that many processes are at work removing oxygen from the air, such as animal life, combustion, decay, the oxidation of mineral matter, and so on, and that at the same time carbon dioxide is being produced in large amount. The principal agencies tending to increase the amount of carbonic acid gas are the first three processes above, but much more is furnished by subterranean sources. It would be reasonable to suppose that, under these conditions, the total oxygen should decrease and the carbon dioxide increase. This is not so, however, for experiments over long periods of time show that the amounts of oxygen and carbon dioxide remain constant, and the explanation is to be found in the compensating or balancing action of green plants under the action of sunlight. These, as we learned in Expt. 114,

absorb the carbon dioxide, decompose it, utilise the carbon, and set free the oxygen once more in the pure state.

Nitrogen in its Relation to Plants and Animals.—

The food consumed by animals contains much nitrogen, and when we examine the initial source of all such foods, we find that they are obtained from plants. (From whence do the plants obtain their nitrogen? Is it from the air?) Our experiments teach us that nitrogen is a very inactive substance—animals cannot utilize it, as Expt. 113 clearly demonstrates—and you will not be surprised to hear that, save for a few plants of the pea tribe—the *Leguminosæ*—plants are unable to utilize the vast store of free nitrogen in the air. The air, especially after thunderstorms, contains small quantities of free ammonia and nitric acid, which, dissolved by the rain and absorbed by the plant-roots, furnish some small store. We find, however, that the soil contains many nitrates, and these are also utilized. Moreover, large quantities of ammonia and nitric acid, etc., are produced by the germs of putrefaction or decay—*i.e.*, by the oxidation of dead matter (See Nitrification, p. 236). Thus, the plant, by absorbing through its roots carbon dioxide and water containing salts, builds up the complex nitrogenous substances necessary to the animal, who, during his life-processes, breaks these down into simpler ones, returning them once more to supply fresh generations of plants. It is thus rendered possible, by this mutual action, for a limited supply of carbon dioxide, water, and nitrogen, to produce countless generations of living matter.

We are now in a position to discuss the chief constituents of the air and their uses.

1. **Oxygen** is the most essential, since it is the only supporter of animal life and combustion.

2. **Nitrogen** owes its importance mainly to its inactivity, for it serves to weaken or dilute with safety the very active oxygen.

3. **Carbon dioxide** is essential to all green plants, whilst without

4. **Water**, life would be impossible.

CHAPTER XII

THE STUDY OF LIMESTONE

HAVING studied water and air, we now turn to the solid matter composing the "earth." We learned in Expt. 53 that **only part** of the soil is soluble in water, and, moreover, that when soil is heated (Exer., p. 42), it decreases in weight and becomes quite a different substance, so that in all probability *earth* is very complex.

One of the commonest of rocks is **limestone**, which occurs in many different forms—*e.g.*, *marble*, *chalk*, *oyster-shell*, etc. The nature of this substance was investigated by **Dr. Black** (1755), whose experiments we repeat below.

EXPT. 115.—To examine the Action of Solvents on Limestone.—Try to dissolve powdered limestone in water, spirit, dilute sulphuric and hydrochloric acids, as in Expts. 25, 26, and observe closely all that occurs. Do you not discover that it is **insoluble in water**, but **dissolves in acids with effervescence**? Is the substance that remains, after evaporating the solution in acids, limestone? Test it with acids and water. (What do you learn?) Our experiment teaches us that limestone is *not a mixture*, whatever else it may be.

EXPT. 116.—To discover if Any Change in Weight occurs during the Solution of Limestone in Acids.—

We will now repeat Dr. Black's experiment. Place about 5 gms. of limestone in a flask, pour in 100 c.cms. of dilute hydrochloric acid, place in the neck of the flask a plug of cotton-wool (Why?), and weigh the whole. When the action is complete, reweigh. Do we not learn that during the solution of limestone in acids something escapes? (What is the objection to this apparatus? How could you avoid it?)

EXPT. 117.—To discover the Effect of Heat on Limestone.—Weigh a clean crucible and place in it 1 gm. of powdered limestone and reweigh. Heat this strongly over a blowpipe flame or in a muffle-furnace for half an hour or so; cool and reweigh. (What do you learn?) Meanwhile, heat some small pieces of limestone in the Bunsen flame (or blowpipe flame preferably), holding them by means of a pair of crucible-tongs until they glow throughout. (What is the residue? Is it limestone?)

EXPT. 118.—To examine the Residue obtained by heating Limestone.—Add water to it, and observe the clouds of steam given off. The mixture becomes hot. (What is this a sign of?) Add more water, filter, and evaporate. (Is it soluble?) Try the effect of acids on the residue. (Is there any effervescence? Does it dissolve?) Repeat Expt. 117. (Is there any change in weight?)

Do we not learn from these experiments that **limestone when heated is completely changed**—it loses its insolubility in water, its power of effervescing with acids—into a new substance which unites with water with great vigour, dissolves in water, is soluble in acids without effervescence, and weighs considerably less than before?—i.e., something has escaped.

EXPT. 119.—To study the Effect of exposing this Residue to the Air.—Place in a dish and weigh about a gram of the substance obtained in Expt. 117. Expose it in the air until the next lesson. Reweigh. (What do you discover?) Test it with acids and water. (Is it, or any of it, limestone again? Does not this experiment teach us that this residue absorbs something from the air and re-forms limestone?) Expose a beaker of a solution of this residue for a week (What do you notice?); collect the residue and test it as above. (Is this limestone?)

EXPT. 120.—To discover what it is that escapes when Limestone is heated.—In the small ignition-

tube (*a*), resting on the retort-stand, place a small quantity of powdered limestone, attach to it a narrow delivery-tube (*b*), and arrange to collect any gas given off in a test-tube (*c*) full of

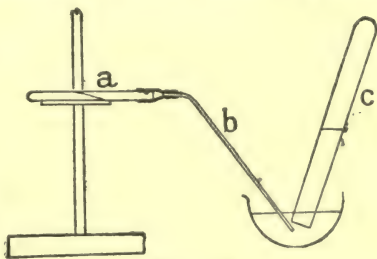


FIG. 56.

water, standing in a small dish of water. Strongly heat the limestone, and observe that (1) a gas collects in *c*, and (2) a white residue is left in *a*. **Limestone is therefore a compound substance.**

We are now in a position to explain what occurs when limestone is heated, for a gas escapes—hence the weight decreases—and leaves an entirely new substance—a strong caustic alkali. (What is the gas?)

EXPT. 121.—To examine the Gas from Limestone.—1. Apply a light to the gas; the gas does not burn, and the taper is extinguished. (Is the gas nitrogen?)

2. Using the apparatus, Fig. 56, bubble the gas through (1) a little clear lime-water and shake well: the lime-water becomes turbid (What gas is it?); (2) a little blue litmus and shake well; it is reddened. **The gas is that produced when a candle is burned in air or charcoal in oxygen—viz., carbonic acid gas.**

Is this the gas liberated by acids in Expt. 116?

EXPT. 122.—To examine the Action of Acids on Limestone.—Place some limestone in the bottle (*a*), fitted as shown (Fig. 24). Pour down the thistle funnel some dilute hydrochloric acid. Effervescence occurs and a gas collects. Test it as above; it proves to be carbonic acid gas, and it is the escape of this gas that causes the loss in weight when limestone is dissolved in acids, as we found in Expt. 116.

The residue obtained when limestone is heated is termed **lime**, or **quicklime** (hence the name *limestone*); and many of you have no doubt seen limekilns in operation. We next proceed—

EXPT. 123.—To compare Lime and Limestone more completely.

1. We have already learned (Expt. 115) that **limestone is insoluble in water, and has no apparent action with water**; whereas **lime (Expt. 118) is soluble in water, and unites with water with great vigour**, steam escapes, and a dry powder is left if too much water is not added. This is known as **slaking lime**, and the product is **slaked lime**.

2. We also learned that **limestone effervesced with acids (Expt. 122)**, giving off carbon dioxide, but that **lime does not (Expt. 118)**, although both neutralize the acid and yield the same salt. Repeat the experiment and verify the latter statement.

3. Place a piece of limestone and lime respectively on separate pieces of wet red and blue litmus-paper. The pieces under the limestone are unaffected—*i.e.*, **limestone is neutral**, whilst the red litmus is turned blue by the lime—*i.e.*, **lime is alkaline**.

4. Into two clean weighed crucibles place about 0.5 gm. of limestone and lime respectively and reweigh. Now heat both strongly for half an hour; cool and reweigh. You will find that **limestone loses weight when it is heated** (What is this due to?), while **lime remains constant in weight**.

5. Expose to the air for a week weighed quantities of limestone and lime in separate dishes; reweigh: the **limestone is constant in weight**, whilst the **lime gains in weight**. This, as we learned in Expt. 119, is due to the fact that it absorbs carbon dioxide (and water vapour) from the air.

6. Place about a gram of limestone and lime in two clean porcelain boats (*a* and *b*), and weigh. Insert these in a hard-glass tube (Fig. 57), through which pass carbonic acid gas for half an hour or so, and then

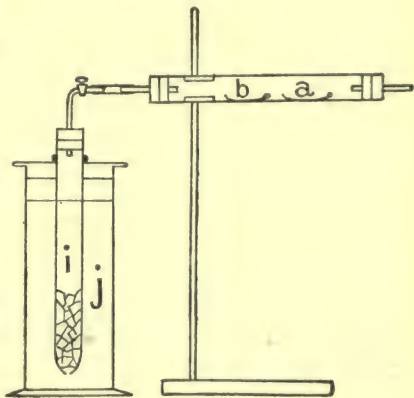


FIG. 57.

reweigh. Observe that the **lime gains in weight** (What has happened?), but the **limestone remains unchanged**.

We may thus conclude that limestone (marble, chalk, etc.) is a compound of lime and carbonic acid

gas; it is therefore termed by chemists **calcium carbonate**, whilst **lime** (quicklime), for reasons you will learn later, is named **calcium oxide**.

There are many other carbonates besides calcium carbonate. Do they all behave in the same way?

EXPT. 124.—To examine the Behaviour of Carbonates when heated.—Heat small quantities of as many carbonates as you can in a small heating-tube. Into a test-tube place some clear lime-water and hold them together mouth to mouth. After heating, shake the test-tube and lime-water and verify the statement that all carbonates (except sodium and potassium) when heated give off carbon dioxide, and leave an oxide of the metal.

EXPT. 125.—To study the Action between Carbonates and Acids.—In a test-tube place a small amount of a carbonate and pour on it some dilute sulphuric acid, holding the tube, as before, mouth to mouth with a tube containing clear lime-water. Repeat this experiment, using various carbonates and various acids, and verify the following *rule*: **All carbonates effervesce with all acids, giving off carbon dioxide; water and a salt are also formed in each case, which may be obtained by filtering and evaporating.**

We will next proceed to investigate, in more detail, the gas which is liberated in the above, and many other, reactions.

CHAPTER XIII

CARBON DIOXIDE (CARBONIC ACID GAS)

CHEMICAL formula, CO_2 . Molecular weight, 44
Density, 22 (or 0.00198 gms. per c.c.).

This gas—the *gas sylvestre* of Van Helmont—is, as we have already seen, present in the atmosphere, being produced, by the respiration of animals, the combustion of carbon compounds, and decay. It is found in caves, deep wells, etc., and is exceedingly plentiful combined with basic oxides such as carbonates (the commonest being calcium carbonate), as limestone, chalk, marble, coral, etc., and as the beautiful stalactites and stalagmites found in limestone caves. It is also formed during alcoholic fermentation, and large quantities of this gas are now brought into the market from breweries, etc.

Methods of making Carbon Dioxide.—In Expts. 74 and 98 we found that when charcoal [A] or compounds of carbon (candle, etc.) are burned in air or oxygen, carbon dioxide results. In Expt. 124 it was discovered that carbonates when heated gave off carbon dioxide [B], and in Expt. 125 that this gas was rapidly set free when an acid decomposed any carbonate [C]. The latter is the most convenient method for preparing any large quantity of the gas in the laboratory.

EXPT. 126.—To prepare and collect Carbon Dioxide.—Into the two-necked—"Woulffie's"—bottle (Fig. 58, *a*) place a quantity of a carbonate (marble is generally used). Fit it with a cork and thistle funnel, and

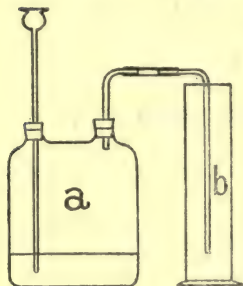


FIG. 58.

a cork and delivery-tube passing into a gas-jar (*b*). Pour dilute hydrochloric acid (any other acid would do) down the thistle funnel and test with a lighted taper to see if the jar is full. At first the light is extinguished at about the middle of the jar, then a quarter of the way down, and finally at the mouth. This method of collection is called by *displacing air upwards*, and can

only be used when the **gas is denser than air**. The salt formed in this case—*calcium chloride* [**C**]*—*may be obtained by filtering the liquid in *a* and evaporating.

EXPT. 127.—To examine the Properties of the Gas.—This gas, as we have already learned, is a **colourless, odourless, faintly-acid-tasting** gas (the sharpness of soda-water, etc., is due to this substance). It is, as we have seen by its collection, **denser than air**; compared with air its density is about 1.5.

1. Light a taper and pour carbon dioxide on it; the light is put out, the gas does not burn; the gas is dense, since it can be poured.

2. Place a little lime-water in the gas-jar containing air and pour carbonic acid gas from a second jar into it. Shake up the gas in the former; the lime-water becomes turbid, showing that carbon dioxide flows downwards, or is denser than air.

3. Pour carbon dioxide from a gas-jar into a beaker,

counterpoised on a balance. The beaker begins to sink, showing the same fact as before.

4. Place a little blue litmus in a jar of the gas and shake well; it is reddened—*i.e.*, the gas is **acidic** (see Expt. 74). Since the gas is an acid, we must dry it with a neutral or an acid-drying agent.

5. **To find the Density of Carbon Dioxide.**—In order to do this correctly, we must first learn how to **prepare pure dry carbon dioxide**. The gas, prepared as in Expt. 126, is dried by passing it through a U-tube containing calcium chloride or strong sulphuric acid, and after allowing a small amount to escape (Why?), collecting it over mercury, as in Fig. 24. In order to determine the density, we proceed as follows: Fill the weighed flask used in Expt. 59 (10) with pure dry carbonic acid gas and weigh. Calculate its density and verify the statement that the **density of carbon dioxide is 0.00198 gm. per c.cm., or twenty-two times that of hydrogen**.

6. Place a jar of air *full* of the gas mouth downward in a trough of water; leave it for some time, and note the level. It continues to rise, showing that the **gas is soluble in water**. Test the solution in the gas-jar—

(1) With blue litmus: it is reddened—*i.e.*, is an **acid**.

(2) With clean lime-water: it is rendered milky—*i.e.*, **carbonic acid [D]**.

(3) Boil a portion vigorously and test the gas by holding a tube with lime-water in the vapour. Test the remaining liquid with clean lime-water. The **gas is driven off carbonic acid by boiling**.

7. Plunge into the gas fragments of burning sulphur, phosphorus, etc., on deflagrating-spoons, and observe that the **gas does not burn, nor does it allow things to burn (or live) in it**. This explains the use of liquid

carbon dioxide as a *fire extinguisher*, and the danger of ill-ventilated rooms, deep wells, cellars, brewers' vats, coal-mines, etc. The deadly *after-damp* of the coal-mine consists largely of this gas.

To examine the Composition of Carbon Dioxide.—This gas, as we learned in Expt. 74, was formed by burning charcoal—*i.e.*, the element carbon—in oxygen. Both the carbon and oxygen disappear, while carbon dioxide is formed. We have thus built up (or *synthesized*) carbon dioxide from the elements carbon and oxygen. The truth of this may also be found by analysis.

EXPT. 128.—To show the Presence of Carbon and Oxygen in Carbon Dioxide.—Into the drawn-out tube (Fig. 30, *a*) place a boat containing a little of the element potassium, and pass carbon dioxide through the apparatus. On heating, the potassium burns, a white residue of **oxide** of potassium being formed (hence **oxygen is present** in the gas, since potassium is an element) and a black layer of **carbon** [**E**]. If a strip of burning magnesium ribbon is plunged into a jar of the gas, it continues to burn, **oxide** of magnesium and **carbon** [**F**] resulting.

EXPT. 129.—To find the Composition by Weight of Carbon Dioxide.—Weigh a clean boat (*d*) containing about $\frac{1}{2}$ gm. of pure charcoal, and place it in the hard-glass tube (Fig. 59) filled at *c* with pure dry oxide of copper to complete the oxidation. Oxygen is passed through the apparatus when all is red-hot, the products being collected in the weighed tubes (*e*, *f*), containing caustic potash or soda lime. When the carbon has burned away and the apparatus is cool, disconnect and reweigh *d*, *e*, and *f*.

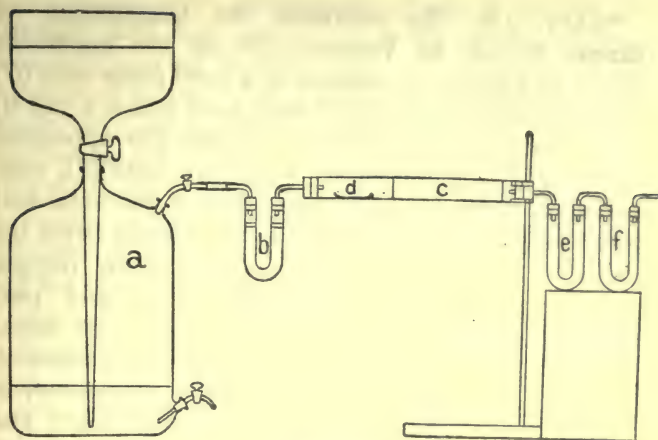


FIG. 59.

The results are calculated thus:

RESULTS—

- | | |
|---|---------------------------|
| (1) Weight of <i>d</i> before | = 1.64 gms. |
| Weight of <i>d</i> after | = 1.17 „ |
| ∴ Loss = weight of carbon | = 0.47 gm. |
| (2) Weight of <i>e</i> and <i>f</i> after | = 47.55 gms. |
| Weight of <i>e</i> and <i>f</i> before | = 45.93 „ |
| ∴ Gain = carbon dioxide | = 1.62 gm. |
| (3) Weight of oxygen | = 1.62 - 0.47 = 1.15 gms. |

I.e.,

1.62 gms. of carbon dioxide contain 1.15 gms. of oxygen;
or 1 gm. of carbon dioxide contains

$$\frac{1.15}{1.62} = 0.71 \text{ gm. of oxygen.}$$

Similarly, 1 gm. of carbon dioxide contains

$$\frac{0.47}{1.62} \text{ or } = 0.29 \text{ gm. of carbon.}$$

***EXPT. 130.—To determine the Composition of Carbon Dioxide by Volume.**—Fit up the apparatus shown in Fig. 60. It consists of a hard-glass tube (*G*) (about 6 inches long), fitted with good corks, a tap (*t*) and a wide-limbed U-tube (*c*). When the apparatus is air-tight, place a small porcelain boat (*d*), containing a small piece of charcoal, into the tube and pour just sufficient mercury in the U-tube to cover the bend.

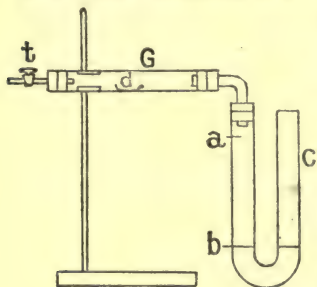


FIG. 60.

Attach an oxygen generator to *t*, and pass oxygen through the apparatus until all air is expelled. Close *t* and pour mercury in the U-tube, opening *t* just a trifle, until the mercury reaches the mark *b*. Turn off *t* and heat the carbon strongly with a *small* flame. The charcoal burns, forming

carbon dioxide [*A*], and on cooling, the mercury is found to be stationary once more at *b*. Thus, carbon dioxide contains its own volume of oxygen. This confirms the result obtained in the preceding experiment, for 1 volume of carbon dioxide weighing 22 parts contains 1 volume of oxygen weighing 16 parts; and hence 1 volume of carbon dioxide contains 6 parts of carbon by weight. Thus, carbon dioxide is made up of $\frac{16}{22}$ —i.e., $\frac{8}{11}$, or 72.72 per cent. of oxygen, and $\frac{6}{22}$; i.e., $\frac{3}{11}$ or 27.27 per cent. of carbon by weight, verifying the results obtained in Expt. 129.

Carbonic Acid.—The solution of carbon dioxide in water is, as we have seen, acid to litmus, but loses this property when boiled

EXPT. 131.—To Examine the Action of Alkalies on Carbonic Acid Gas.—Shake up in test-tubes of carbon dioxide, covering the end with your thumb, (1) some solution of caustic soda and caustic potash, (2) lime-water, (3) a piece of damp, solid caustic soda or potash. (Is the gas absorbed?) Open the tubes under water (What do you observe?) Add a spot of dilute acid to each. (What happens? What substances are produced?) We thus learn that **alkalies readily absorb the gas forming the salts, carbonates.**

***EXPT. 132.—To find the Volume of Carbon Dioxide in a Gaseous Mixture**, in which the only acid gas is carbon dioxide—*e.g.*, oxygen, nitrogen, and carbon dioxide—we make use of the fact discovered in the last experiment.

1. Fill a graduated tube standing over mercury with the mixture. Insert into the gas a small piece of damp caustic soda or potash on the end of a piece of copper wire, and observe the level of the mercury. When the mercury ceases to rise, remove the caustic soda and transfer the tube, by placing a crucible under the open end, into a large cylinder of water as shown (Fig. 61); level the water inside and out, and read the volume.

2. Expt. 109 may be repeated, the tubes *a* and *b* being filled with mercury, the pipette (*p*) with a strong solution of caustic soda.

N.B.—The oxygen in the residual gas could then be determined as in that experiment.

We discovered, by repeating Dr. Black's experi-

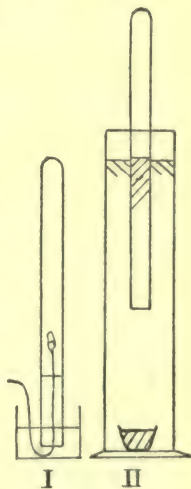


FIG. 61.

ments, that limestone or calcium carbonate, when strongly heated, was converted into a strongly **caustic alkali lime** (**Calcium oxide [B]**). This we found was a white powder, which unites with vigour to form a dry white powder—**slaked lime**—clouds of steam escaping; and, further, that this slaked lime dissolved in more water, forming an alkaline solution—**lime-water**—which we have used as a test for carbon dioxide. We also discovered that it slowly absorbs carbon dioxide, and re-forms calcium carbonate (see Expt. 123). Let us examine the action of carbon dioxide a little more closely.

EXPT. 133.—To study the Action of Carbon Dioxide on Lime-Water.—Pass a current of carbon dioxide through half a beaker full of clear lime-water. A heavy white *precipitate* (Why so named?) is formed. Divide the milky liquid into two parts and proceed thus:

First Portion.—Filter and wash the precipitate on the filter-paper, and to the insoluble add a spot of acid; it effervesces—it is a **carbonate of lime [G]**.

Second Portion.—Continue to pass carbon dioxide until this insoluble precipitate disappears, and test as follows: (1) Add a spot of acid; effervescence occurs—it is a **soluble carbonate of lime [H]**. (2) Boil a second portion, notice the effect of the escaping gases on lime-water (What escapes?), and observe that a precipitate **[I]** is formed; test this with acid. Is this the same as the residue obtained in the first portion? (3) To a third portion add an equal bulk of fresh lime-water and observe the effect. Test the product with acid.

It seems, then, that lime forms **two** carbonates: (1) **Insoluble**, produced with a **smaller** amount of carbon dioxide. This was studied in Expt. 123, and

was found to be **neutral calcium carbonate**. (2) A **soluble** one, produced with a **larger** amount of carbonic acid, which readily loses carbon dioxide on being heated, forms the neutral salt, and can neutralize more base; this is the **acid salt, calcium bicarbonate**.

EXPT. 134.—To examine the Action of Carbon Dioxide on Caustic Soda Solution.—Caustic soda, like lime, absorbs carbon dioxide. (Does its solution behave in the same way?) Repeat Expt. 133, testing from time to time with red litmus. Does any solid separate? Does the solution even become neutral? The action, so far as appearances can tell us, is quite different from that of lime-water. (Why is this?)

EXPT. 135.—To examine the Solubility of Carbonates.—Shake up small quantities of as many carbonates as you can with water; filter and evaporate to dryness and verify the statement that carbonates (except **potassium sodium and ammonium**) are insoluble in water.

EXPT. 136.—To examine the Action of Carbonates with Litmus.—Shake up small quantities of various carbonates with red and blue litmus solution and verify—

1. **Insoluble carbonates are neutral.**

2. **Soluble carbonates are either** (1) like potassium of sodium carbonates, **alkaline**, or (2) like calcium bicarbonate, **acid** to litmus.

Hence the difficulty met with in Expt. 134. Now we have discovered that bicarbonates when heated or boiled (see Expt. 133) lose carbon dioxide and water, and form the **normal carbonates**; hence

EXPT. 137.—To prepare Soluble Normal Carbonates.
—We proceed thus: Pass carbon dioxide through a strong solution of an alkali for a considerable time. A mixture of the normal [**J**] and acid [**K**] carbonates is thus formed: (1) Boil the liquid; the acid carbonate [**L**] is destroyed; cool and crystallize; or (2) evaporate the liquid to dryness and heat strongly. Dissolve the residue in water and crystallize.

EXPT. 138.—To examine Sodium Carbonate (Washing-Soda, Soda Crystals).—This is a white crystalline solid. Expose a clean crystal to the air; it crumbles to a dry, shapeless powder—*i.e.*, it is efflorescent. Shake up a small crystal with water—it dissolves; add red litmus—it turns blue—*i.e.*, it is alkaline. In Expt. 124 we learned that it is not decomposed by heat, but in Expt. 125 that it effervesces with acids, liberating carbon dioxide and forming sodium salts and water. Washing-soda contains, as we learned in Expt. 50, about two-thirds of its own weight of water of crystallization.

EXPT. 139. — To prepare Sodium Carbonate. —
1. Bubble carbon dioxide through a solution of sodium hydrate until the gas freely escapes; boil the liquid to a small bulk, cool it, and crystallize.

2. Heat a weighed quantity of bicarbonate [**L**] of soda in a crucible; test the gas evolved with a little lime-water on a glass rod: it is carbonic acid. Test the gas evolved by holding a cold dish over the crucible; moisture condenses; cool and weigh, and repeat until of constant weight; dry sodium carbonate remains. (If crystals are required, how would you get them from this ?)

3. Add some sodium hydrate to twice its weight of sodium bicarbonate [**M**]; dissolve in water, evaporate, and crystallize.

EXPT. 140.—To examine Sodium Hydrogen Carbonate (Baking-Soda, Bicarbonate of Soda).—This is a white powder. Shake a little with water; it dissolves, but is much less soluble than washing-soda. Test the solution with litmus; it is alkaline. As in the experiment above, when heated it loses water and carbon dioxide (hence its name, “baking-soda,” for in baking the escaping gas renders the bread light and porous), and forms sodium carbonate. Like all carbonates (see Expt. 125), it is decomposed by acids liberating carbon dioxide and forming sodium salts.

EXPT. 141.—To prepare Sodium Bicarbonate.—1. Make a saturated solution of caustic soda, and when cold, bubble carbon dioxide through until no more is dissolved. A white powder is precipitated. Filter and wash with a *very little* water, dry by pressure between filter-paper and exposure to the air. Examine it as above; it is sodium bicarbonate [**K**].

2. Make a saturated solution of washing-soda [**N**], and repeat the experiment. Sodium bicarbonate results.

Potassium carbonate and bicarbonate are prepared like the corresponding sodium compounds [**O**], and have very similar properties.

As, however (see Expt. 135), most carbonates are insoluble, they must be prepared in a different manner. Let us see if we can re-form limestone from lime in any way other than those of Expts. 123 and 133.

EXPT. 142.—To study the Behaviour of Lime on Soluble Carbonates.—1. Dissolve seven parts of washing-soda in water in a beaker, and add to it one part of lime. Heat the mixture nearly to boiling for half an hour or so, stirring frequently. Allow the mixture to settle, and remove a spot of the clear liquid

with a glass tube and drop it into some dilute acid. (Is it washing-soda?) If any effervescence occurs, reheat until, on testing, the liquid is free from carbonate. (How will you know?) Decant the clear liquid and evaporate it to dryness. (What is it?) Test it with litmus; it is alkaline. Feel it; it is soapy to the touch. Expose a little to the air; it deliquesces. Expose a solution of this solid; no lime is deposited—it is, in fact, **caustic soda** [P]. Wash the residue in the beaker frequently with water and test the washed solid with acid. (Is it lime? What is it?)

Repeat this experiment with potassium carbonate [Q]; caustic potash and limestone result. Thus, **lime renders these carbonates into “caustic” alkalies**, itself uniting with the carbon dioxide to form limestone.

2. Add a solution of washing-soda to some lime-water. (What do you observe?) Collect and wash the precipitate, and test it as before; again it is limestone [P]. (What does the filtrate contain?)

3. Dissolve some of the lime in dilute hydrochloric or nitric acid, and to the solution add potassium carbonate until no more precipitate is produced. Boil the mixture and collect the precipitate on a filter-paper. Wash well and test the solid; it proves to be calcium carbonate [R], or limestone. Evaporate the solution to a small bulk and cool. Crystals of potassium chloride, or nitrate, are formed.

These are examples of the methods used for obtaining insoluble carbonates.

EXPT. 143.—To prepare Insoluble Carbonates by Double Decomposition.—To a solution of copper sulphate [S] add a solution of sodium carbonate until, on stirring and allowing the mixture to stand, the liquid

is colourless. Filter the mixture and evaporate the clear filtrate to a small bulk, and set it aside to crystallize. Colourless crystals, quite different from those of washing-soda, are formed. They are crystals of sulphate of sodium. Wash the insoluble residue and test it with acid; this blue solid is copper carbonate. Repeat this with other solutions—*e.g.*, calcium chloride, zinc sulphate, etc. In such experiments **two substances are decomposed and two new bodies produced** by the metals and acids each changing partners. Such a change is called **double decomposition**.

Other Common Carbonates.—We have already studied sodium carbonate and sodium bicarbonate. Potassium, a metal very closely related to sodium, forms similar compounds, which are prepared in exactly the same way from caustic potash.

Potassium Carbonate is present in wood ashes; hence the name “potash.” It is a white, very soluble solid, deliquescent, alkaline to litmus, and neutralizes all acids. It is unaffected by heat, and precipitates the insoluble carbonates of the metals.

Potassium Hydric Carbonate is less soluble than the normal salt. It is decomposed by heat, carbon dioxide and water escaping, forming the normal carbonate.

Magnesium Carbonate is a bulky, white, insoluble solid, decomposed by heat and dilute acid. It occurs native as magnesite, and with calcium carbonate as dolomite. When made by precipitation (Expt. 143), it contains magnesium hydrate.

Zinc Carbonate is a white, insoluble solid, decomposed by heat and dilute acids. It occurs in nature as *calamine*, and when prepared by precipitation, contains zinc oxide.

Lead Carbonate occurs in nature as cerussite. *White lead* is a basic carbonate of lead, which is a valuable pigment.

We are now in a position to explain what is understood when we speak of **hardness of water**. It is well known that some natural waters readily give a good lather with soap, whereas others merely give a thick scum (*curd*), with little or no lather. The former

are called **soft waters** (rain-water is often called “soft water”), the latter **hard**.

EXPT. 144.—To discover the Cause of Hardness.—Shake up in a clean stoppered bottle 20 c.cms. of distilled water with soap solution (made by dissolving soap in a mixture of water and methylated spirit), added 1 c.cm. at a time. When a good lather which is permanent for five minutes is obtained, note the volume of soap solution used. Now repeat, using 20 c.cms. of distilled water, first adding a few grains of various soluble solids—*e.g.*, common salt, sugar, copper sulphate, ferric chloride, etc. You will find that when *salts* are added, the water is rendered hard, and moreover, in the last two cases when a lather is produced, that the solution is colourless, while the curd is coloured.

Does not this prove that **hardness of water is produced by salts in solution**, and the waste of soap is due to its being consumed to precipitate the dissolved solids? (Why is rain-water softer than spring or river water?)

As we learned in Chapter V., the solids in solution in a natural water depend upon the life-history of the water. Water which flows over hard, insoluble rock (*e.g.*, granite) contains very little solid matter, and is soft, whilst water which flows over the softer soluble rocks will be hard. Now it is often found in boilers and kettles that a *crust* or *fur*—a brownish, hard solid—is formed after long use. (What is this composed of?)

EXPT. 145.—To examine the “Fur” of a Kettle.—Take a little *fur* or *boiler crust*, and add a spot of dilute acid to it; effervescence occurs. Test the gas by bubbling some through lime-water. The gas is carbon dioxide and the fur is a carbonate. Further

experiments prove it to be impure calcium carbonate. Now, in Expt. 115 we learned that calcium carbonate is insoluble. How, then, is it found in water ?

EXPT. 146.—To show that Chalk is dissolved by Water containing Carbonic Acid Gas.—Place some chalk in distilled water in a beaker and bubble carbon dioxide through the mixture for an hour or two. Filter the liquid and examine the clear liquid, thus:

1. Add a spot of acid; the liquid effervesces.
2. Add a little fresh lime-water [T]; a white precipitate is formed.
3. Boil a portion; chalk [I] or fur is deposited, whilst carbonic acid gas escapes.
4. Shake up a portion of the clear liquid with soap solution and a portion of the clear liquid obtained in (3) after filtration. (Which is the softer ?)

Compare your results with those of Expt. 144. Do we not learn that some hardness is caused when water containing carbon dioxide dissolves chalk in the form of calcium bicarbonate [H], and that this hardness can be removed by boiling (therefore called **temporary hardness**), since this destroys the soluble bicarbonate and precipitates the insoluble calcium carbonate ?

EXPT. 147.—To show Permanent Hardness—i.e., hardness which is not removed by boiling.—Take 100 c.cms. of water and dissolve in it a few grains of calcium chloride. Test 25 c.cms. of this with soap solution and note your result; boil the remainder for half an hour. (Is anything deposited ?) Make up the liquid when cold to 75 c.cms. (Why ?), and test 25 c.cms. of this with soap solution. Compare your results before and after boiling. Thus a second kind of hardness (viz., permanent hardness) is due to dissolved salts.

which are not decomposed by boiling. (How could you remove this hardness ?)

How to remove Hardness.—Since “hardness” causes so large a waste of soap and gives considerable trouble commercially, by blocking up steam-pipes, etc., it is of importance to find some method of removing it—*i.e.*, of softening water.

Soap, as we learned in Expt. 144, removes **all** hardness by precipitating the soluble matter as “curd,” but this is very wasteful. **Distillation**, since it removes the water from the all-dissolved solids, will also remove **all** hardness. We learned in Expt. 143 that sodium carbonate will precipitate insoluble carbonates of the metals in solution; and hence in laundries **washing-soda is added [S]**, to remove the permanent hardness and the mixture is **boiled [I]**, which destroys the temporary hardness. (Is all hardness removed ?) Temporary hardness can also be removed by adding lime (see Expt. 133); so that if a mixture of **lime and washing-soda** is added in the proper portion, **all** will be removed. Before we can use this, however, it is necessary to know how much of each hardness is present. (Why ?)

EXPT. 148.—To estimate the Hardness of a Given Sample of Water.—A standard solution of hard water will be required. This is made by dissolving 1.11 gms. of calcium chloride in a winchester—*i.e.*, $2\frac{1}{2}$ litres—of water. Then each c.cm. of this solution is equal to 0.0004 gm. of calcium carbonate. Fill a burette with soap solution and put 50 c.cms. of distilled water in a clean stoppered bottle and run in soap solution, 1 c.cm. at a time, shaking well until a permanent lather is obtained. Repeat this, using 50 c.cms. of the solution, and work out your results as below :

50 c.cms. of distilled water require 3.5 c.cms. of soap solution;

50 c.cms. of standard solution require 24.0 c.cms. of soap solution

—i.e., hardness in 50 c.cms. of standard solution require $24.0 - 3.5 = 20.5$ c.cms. soap.

1 c.cm. of soap solution removes the hardness in $\frac{50}{20.5}$ c.cms. of standard solution,

or 1 c.cm. of soap solution removes $\frac{50}{20.5} \times 0.0004$ gm.
or 0.00098 gm. of calcium carbonate.

This is called **standardizing** the soap solution.

A. To find the Total Hardness of a Given Sample.—

Take 50 c.cms. of the water and find the amount of soap solution required to give a lather—e.g., 34.5 c.cms. of soap solution required—i.e., $34.5 - 3.5$, or 31 c.cms., of soap solution required for the hardness present. Then 50 c.cms. of this water contain hardness equal to that of 31×0.00098 , or 0.0304 gm. of calcium carbonate.

B. To find the Temporary and Permanent Hardness in a Hard Water.—Take 50 c.cms. of the sample and

boil for twenty minutes or so; cool and make up its volume to 50 c.cms. with distilled water. Find the amount of soap solution required for the hardness left after boiling. (Which is this ?)

E.g., before boiling, soap required = 34.5 c.cms.; total hardness requires 31 c.cms. of soap.

After boiling, soap required = 17.5 c.cms.; permanent hardness required 14 c.cms. of soap; whence temporary = 17 c.cms.

I.e., temporary hardness is $\frac{17}{31}$ and permanent $\frac{14}{31}$ of the total.

C. To compare the Hardness of Two Given Samples.—Take 50 c.cms. of each, and find the soap solution required to give a lather with each—*e.g.*,

50 c.cms. of A required 76 c.cms. of soap—*i.e.*, hardness = 72·5 c.cms.

50 c.cms. of B required 34 c.cms. of soap—*i.e.*, hardness = 30·5 c.cms.

$$\frac{\text{Hardness of A}}{\text{Hardness of B}} = \frac{72\cdot5}{30\cdot5} = \frac{2\cdot4}{1} \text{ (approx.)}$$

Quantitative Experiments with Carbonates.

In Expt. 124 we learned that most carbonates, when heated strongly, lose their carbon dioxide and leave an oxide. Hence—

EXPT. 149.—To find the Weight of Carbon Dioxide (liberated by Heat) in Any Weight (say 100 Gms.) of a Carbonate.—Weigh a crucible (no lid is required) (Why?); place in it about a gram of a carbonate and reweigh. Ignite strongly on a pipe-clay triangle and tripod stand for half an hour. Cool and weigh and reheat until, on cooling and reweighing, the weight is constant. (What does this tell you?)

RESULTS—

Weight of crucible and carbonate	= 25·93 gms,
Weight of crucible alone	= 25·31 „

∴ Weight of carbonate = 0·62 gm.

Weight of crucible, etc., before	= 25·93 gms.
Constant weight of crucible, etc., after	= 25·73 „

∴ Loss = carbon dioxide = 0·20 gm.

Therefore 0.62 gm. of carbonate contains 0.20 gm. of carbon dioxide,

1 gm. of carbonate contains $\frac{0.20}{0.62}$ gm. of carbon dioxide,
and 100 gms. of carbonate contain

$$\frac{0.20}{0.62} \times 100 = 32.19 \text{ gms. of carbon dioxide.}$$

EXERCISES.—Try zinc carbonate, magnesium carbonate, calcium carbonate, lead carbonate, etc.

In Expt. 125 we found that all carbonates are decomposed by acids losing carbon dioxide, and we may use this for the same purpose.

EXPT. 150.—To ascertain the Weight of Carbon Dioxide (liberated by Acid) in Any Weight (say 100 Gms.) of a Carbonate.—Fit up the apparatus shown in Fig. 62. It consists of a 2-ounce flask (*b*) fitted with a tube (*c*), bent as shown, and having a small side-hole at *d*. Attached to this is a small drying-tube (*e*), containing *fine, freshly heated*, calcium chloride, between two plugs of cotton-wool. Weigh out about 2 gms. of the carbonate on glazed paper; transfer this carefully to the flask. Wash all the solid from the neck of the flask and half fill the flask with water. Attach the small tube *t* to the tube *c* at *d* by means of cotton, and arrange, by withdrawing *c* through the cork, that *t* is well above the liquid. Fill *t* with strong nitric acid, replace in the flask, and weigh the whole apparatus. By gently lowering *c*, bring the acid *slowly* (Why?) into contact with the carbonate, and shake the flask very gently. Decomposition ensues, and dioxide escapes through *e*, being dried there. (Why is this necessary?) If the action is too rapid, raise *t*,

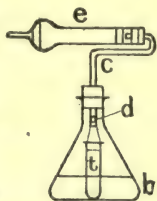


FIG. 62.

and as it slows, lower it more and more until all the carbonate is decomposed, and reweigh.

RESULTS—

Weight of apparatus before = 95.43 gms.

Wgt. of apparatus without carbonate = 91.89 „

∴ Weight of carbonate = 3.54 „

Weight of apparatus before = 95.43 gms.

Weight of apparatus after = 93.87 „

∴ Loss = carbon dioxide = 1.56 „

So that 3.54 gms. of carbonate liberate 1.56 gms. of carbon dioxide;

1 gm. of carbonate liberates

$$\frac{1.56}{3.54} \text{ gms. of carbon dioxide,}$$

and 100 gms. of carbonate liberate

$$\frac{1.56}{3.54} \times 100 = 44.07 \text{ gms. of carbon dioxide.}$$

What sources of error are there in this experiment? How would you obviate them? Repeat Dr. Black's method of conducting the above experiment (see experiment and compare your results). Explain any difference that you find.

EXPT. 151.—To determine the Volume (and Weight) of Carbon Dioxide given off from Any Weight (say 100 Gms.) of a Carbonate when Heated.—Fit up the apparatus shown in Fig. 34. Weigh the tube (a); introduce about 1 gm. of the carbonate and reweigh. Fit up the apparatus as shown, and test to see if it is air-tight. Before commencing to heat, see that the aspirator and exit-tube are filled with water. Now heat strongly and collect the water expelled in the jar (c). When the level of the water in c ceases to rise, cool (Why?) and observe the level. (Why does this descend?) When no further lowering occurs, disconnect and empty

the delivery-tube *c* into the jar. Measure the water in the jar *c*; this is the volume of the gas liberated (What error is there in this experiment? How would you avoid it?)

RESULTS—

Weight of tube and carbonate = 2.29 gms.

Weight of tube alone = 1.24 „

∴ Weight of carbonate = 1.05 gms.

Volume of gas = 232 c.cms.

So that 1.05 gms. of carbonate liberate 232 c.cms. of gas;

1 gm. of carbonate liberates $\frac{232}{1.05}$ c.cms. of gas,

and 100 gms. of carbonate liberate

$$\frac{232}{1.05} \times 100 = 22,095 \text{ c.cms. of carbon dioxide.}$$

Why is it unnecessary to use more than 1 gm. of carbonate in this experiment?

By weighing the tube (*a*) when cold, the loss in weight—*i.e.*, the weight of carbon dioxide expelled—can be found.

EXPT. 152.—To determine the Volume of Carbon Dioxide liberated by Acid from Any Weight (say 100 Gms.) of a Carbonate.—Fit up the apparatus shown (Fig. 63), consisting of the small flask (*b*) used

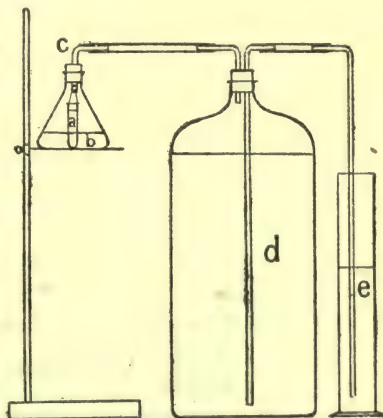


FIG. 63.

in Expt. 150, provided with a cork carrying a right-angled glass tube (*c*), having a side-hole to which

is attached the small tube (*a*). The tube (*c*) is attached to the aspirator, as shown, and when the whole is air-tight, proceed thus: Weigh the small flask; place in it about 2 gms. of the carbonate and reweigh. Wash all the solid into the flask and half fill it with water. Fill the tube (*a*) with strong nitric acid, and, keeping it away from the solid, fit up the apparatus once more. When air-tight, lower *a* and tilt the flask, and the action commences. When all the carbonate is decomposed and the apparatus is quite cold, measure the water expelled as before and calculate as above. (Can you discover any defects in the above apparatus? If so, how would you avoid them?)

In Expts. 150 and 152 we found the weight and volume of carbon dioxide liberated from carbonate by acids, and in Expts. 149 and 151 by the action of heat. Thus, if we use the same carbonate in each of these pairs of experiments, we are obviously measuring the same quantity of gas. Hence from these two sets of experiments we *may deduce the density of carbon dioxide—e.g.*,

In Expt. 150 100 gms. of zinc carbonate yielded
42.15 gms. of gas;

In Expt. 152 100 gms. of zinc carbonate yielded
22,095 c.cms. of gas;

So that 22,095 c.cms. of carbon dioxide weigh
42.15 gms.;

1 c.cm. of carbon dioxide weighs

$$\frac{42.15}{22,095} = 0.0019 \text{ gm. per c.cm.}$$

Thus, the density of carbon dioxide is 0.0019 gm. per c.cm.

Compare your result with that of Expt. 127.

CHAPTER XIV**

CARBON MONOXIDE

WE have already discovered that carbon burns in air and oxygen, producing carbon dioxide, and the question may arise, Is this the only oxide of this element? When carbonic acid gas is vigorously reduced (see Expt. 128), carbon results. Now, carbon itself, as we learned in Expt. 81, is, when hot, a strong reducing agent, and it would seem reasonable to suppose that if it were possible to reduce carbon dioxide by means of carbon, a lower oxide of carbon would be produced. Let us see if this is the case.

EXPT. 153.—To study the Action of Carbon on Carbon Dioxide.—Pass carbon dioxide slowly through the hard-glass or iron tube (Fig. 27), packed with small pieces of charcoal made red-hot. The issuing gases must first be passed through a wash-bottle containing a strong solution of caustic potash or soda solution (Why?) and any other gas collected over water. Observe that a **colourless gas, insoluble in water**, collects. Test this gas—(1) by shaking with lime-water—no milkiness is produced; (2) by applying a light—it **burns** with a lambent **blue flame**; (3) after burning again, shake well—the lime-water now becomes turbid. Thus a gas has been obtained by **reducing carbon dioxide with carbon [A]**, which burns, producing carbon dioxide—it is **carbon monoxide**.

The gas is also produced by reducing carbon dioxide with red-hot iron.

Steam may be reduced by hot carbon [B] in a similar way, an iron tube being used, and a mixture of hydrogen and carbon monoxide is obtained, known as "water gas." This is much used in manufactures where a smokeless fuel is required.

It was a common practice on the Continent at one time to use braziers of red-hot charcoal for warming bedrooms, etc., and many somewhat mysterious cases of "death from poisoning" occurred. This was discovered to be due to the production of this deadly gas. If we examine a glowing fire closely, although few or no yellow luminous flames can be observed, a curious lambent, blue flame may be noticed playing over the surface of the red-hot coal. These are flames of burning carbon monoxide, produced by the reduction of carbonic acid gas and water-vapour as they pass through the red-hot coal.

Carbon Monoxide (Carbonic Oxide).

Chemical formula, CO. Molecular weight, 28 (= 12 + 16). Density, 14 (or 0.00126 gm. per c.cm.).

This deadly poisonous gas occurs in the *after-damp* following colliery explosions, and in the fumes from lime-kilns, etc. It may be prepared, as above, by the reduction of carbon dioxide by means of carbon and iron or of steam with carbon, but is more conveniently obtained as below.

EXPT. 154.—To examine the Action of Strong Sulphuric Acid on Formates.—Add cold strong sulphuric acid to as many formates as you can in separate test-tubes; *warm*, if necessary, and apply a light to any gas

evolved. Observe that in all cases a gas is produced that burns with the characteristic flame of carbon monoxide. Thus, **all formates are decomposed by strong sulphuric acid setting free carbon monoxide, sulphates of the metals being also produced.**

EXPT. 155.—**To examine the Action of Strong Sulphuric Acid on Oxalic Acid and Oxalates.**—Repeat Expt. 154, using oxalic acid and as many oxalates as you can. Observe that *warming is necessary*, and that the gas burns with difficulty. (Why?) Hold a small tube, with a spot of lime-water on the end, in the gas, and notice that it becomes turbid. Do we not thus learn that **oxalic acid and oxalates are decomposed by strong sulphuric acid on warming, a mixture of carbon dioxide and monoxide being liberated?** (What else is produced?)

EXPT. 156. — **To prepare and collect Carbonic Oxide.** —

1. Place in the flask (a) some **sodium formate [C]** (any other formate could be used), and arrange to collect the gas over water. Cover the formate with strong sulphuric acid, and gently, as before, warm. Examine what collects.

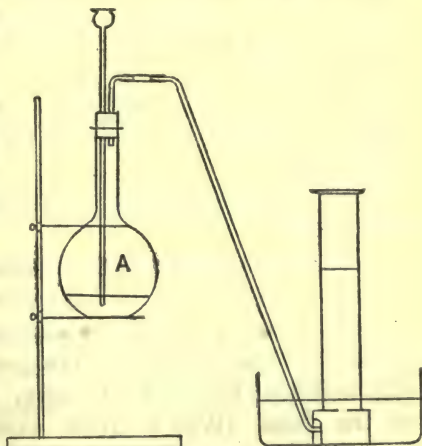


FIG. 64.

2. **Oxalic Acid [D]** is generally employed for the preparation of this gas (oxalates might also be used). Since, as we learned in Expt. 155, a mixture of gases

is given off, it is necessary to pass the escaping gases through a wash-bottle of caustic soda solution before collecting the carbon monoxide.

Larger quantities of the pure gas may be obtained by warming potassium ferrocyanide with sulphuric acid (three volumes of strong acid to one volume of water). Try this in a test-tube.

EXPT. 157.—To examine the Properties of Carbon Monoxide—Caution.—Since the gas is such a **deadly poison**, the greatest care must be taken to avoid inhaling it. We have already learned that it is a **colourless, odourless, tasteless gas, insoluble in water**. It is a **combustible gas**, and readily burns with a *blue* [E] flame, producing carbon dioxide. The gas is without effect on lime-water.

1. Plunge a lighted taper into a jar of the gas; the light is extinguished—*i.e.*, the gas is a **non-supporter of combustion**.

2. Shake up some neutral litmus with the gas—it remains unaffected, so that carbon monoxide is a **neutral gas**.

3. Since the gas burns so readily, it is a **good reducing agent**.

To show that Carbonic Oxide reduces Metallic Oxides.—Repeat Expt. 59 (ii.), placing iron oxide in the hard-glass tube, and passing carbonic oxide through the apparatus. When all the air is expelled, light the issuing gas, and heat the oxide strongly. Note the extinction of the flame (Why? How would you verify your answer?), and examine the residue after cooling in the gas. (Why?) It is found to be metallic iron. Hence, **carbon monoxide reduces iron oxide to iron, being itself oxidized to carbon dioxide** [F]. Repeat this with other metallic oxides—*e.g.*, copper oxide, lead oxide, etc.

4. Introduce into a jar of the gas, standing over mercury, some solution of cuprous chloride, and observe that it is slowly absorbed. This is useful for *removing and determining the amount of carbon monoxide in a gaseous mixture.*

5. Repeat Expt. 59 (10), and verify the statement that the **density of carbonic oxide is 0.00126 gm. per c.cm.** Hence, compared with hydrogen, its **relative density is $\frac{0.00126}{0.00009}$, or 14.**

To examine the Composition of Carbonic Oxide.—Since carbon monoxide is produced from carbon dioxide and carbon entirely, and when burned produces carbon dioxide only, it must be **composed of carbon and oxygen only.**

***EXPT. 158.—To find the Composition of Carbonic Oxide by Volume.**—Repeat Expt. 89, using carbon monoxide and excess of oxygen. After reading the volume after explosion, introduce a little strong caustic soda solution or a piece of the damp solid. When the absorption of carbon dioxide is complete (How will you know?), level the pressure and take the reading again.

RESULTS—

Volume of carbon monoxide = 3.10 c.cms.

Volume of carbon monoxide and oxygen = 20.70 c.cms.

After explosion, volume of carbon dioxide and oxygen = 19.15 c.cms.

After absorption of carbon dioxide, the volume of oxygen = 16.05 c.cms.

∴ Volume of carbon dioxide = $19.15 - 16.05 = 3.1$ c.cms.,
and volume of oxygen used = $(20.70 - 3.10) - 16.05$
= 1.55 c.cms.

So that 3.1 c.cms. of carbon monoxide and 1.55 c.cms. of oxygen produce 3.1 c.cms. of carbon dioxide, or two volumes of carbonic oxide and one volume of oxygen form two volumes of carbon dioxide. This volumetric result may be used to find the composition of the gas by weight; thus—

Two volumes of carbon dioxide contain two volumes of oxygen (see Expt. 130);

∴ Two volumes of carbon monoxide contain one volume of oxygen;

or, carbon monoxide contains half its volume of oxygen.

The densities of these gases (Expts. 157 and 76) are 14 and 16 respectively. Therefore, two volumes of carbon monoxide (weighing 2×18 —i.e., 28 parts by weight) contain one volume of oxygen (weighing 16 parts by weight): so they also contain carbon which weighs $28 - 16$, or 12 parts. So we find that carbon monoxide consists of

$\frac{12}{28}$, or $\frac{3}{7}$ (42.86 per cent.) of carbon, by weight,
and
 $\frac{16}{28}$, or $\frac{4}{7}$ (57.14 per cent.) of oxygen, by weight.

This result may be confirmed thus:

EXPT. 159.—To determine the Composition of Carbon Monoxide by Weight.—Pack the hard-glass tube (Fig. 59) with pure dry copper oxide [G], and after weighing, pass through it *slowly* some pure dry carbonic oxide. When all the air is expelled, heat the oxide strongly, and collect the carbon dioxide formed in the weighed tubes containing soda lime, etc.

RESULTS—

Weight of copper oxide tube before = 45.62 gms.

Weight of copper oxide tube after = 44.60 „

∴ Loss = weight of oxygen = 1.02 gms.

Weight of U-tube after = 24.57 gms.

Weight of U-tube before = 21.76 „

∴ Gain = weight of carbon dioxide = 2.81 gms.

Thus wgt. of carbon monoxide = $2.81 - 1.02 = 1.79$ gms.

Since the weight of oxygen in 2.81 gms. of carbon

dioxide = $\frac{8}{11} (2.81) = 2.04$ gms.,

∴ the weight of oxygen in 1.79 gms. of carbon monoxide is $2.04 - 1.02 = 1.02$ gms.;

Whence the weight of carbon in 1.79 gms. of carbon monoxide is $1.79 - 1.02 = 0.77$ gm.

So that carbon monoxide consists of

$\frac{0.77}{1.79}$, or 43.02 per cent. of carbon, by weight,

and

$\frac{1.02}{1.79}$, or 56.98 per cent. of oxygen, by weight,

thus *confirming* the results obtained in the preceding experiment.

CHAPTER XV

CARBON

CHEMICAL symbol, C. Atomic weight, 12.

This element is familiar to all, especially in its impure forms, coal, coke, charcoal, soot, etc. It also occurs almost pure in the crystalline forms, **graphite** (plumbago, blacklead) and **diamond**. Carbon is widely distributed when combined with other elements, being present in the tissues of plants and animals; it exists in great quantity as carbonates of lime, etc., as limestone, marble, chalk, coral, and many shells, and is an important constituent of the mineral oils—*e.g.*, petroleum, etc.

EXPT. 160.—**To examine the Effect of heating Animal and Vegetable Substances.**—Heat small quantities of flour, starch, sawdust, meat, bread, sugar, etc., first in small tubes, and secondly in iron spoons. Observe that a black, shapeless mass is obtained in the tubes, which burns away in the spoon; it is charcoal or carbon.

EXPT. 161.—**The Destructive Distillation of Wood.**—Fit up the apparatus shown (Fig. 65); *a* is a test-tube about half full of sawdust or broken wood-spills, fitted with a cork and tube leading to the bottle (*b*), provided with an escape-tube. Heat the sawdust strongly,

observe closely all that occurs, and verify the statement that (1) a gas of strong odour is given off, which burns when a light is applied.

Caution.—Wait until all air is expelled from *b* before lighting. (Why?)

(2) Two liquids collect in *b*, one a watery liquid, which is acid to litmus (this is due to acetic acid, which is called **pyroligneous acid**) and the other a **tarry** liquid.

(3) A porous black residue is found in *a*, which is **wood charcoal**. (Why is this termed “destructive disillation”?)

It is obvious that, since coal is formed by nature from wood, we should expect similar products when it is distilled. Repeat the experiment, using coal instead of sawdust. **Coal-gas** is given off, **coal-tar** and an aqueous liquid is produced, and **coke** remains.

(Can you explain what occurs in an ordinary coal fire?)

Amorphous—*i.e.*, **non-crystalline**—**carbon** is familiar to all in the form of **soot**. When prepared by igniting resin, camphor, etc., away from the air, it is termed **lampblack**, which is moderately pure. An impure form of **charcoal** is prepared as above by destructive distillation from wood, whilst coal similarly treated yields **gas carbon**—a hard, pure form—and **coke**—a very impure variety. **Animal charcoal**, or **bone-black**, is made by distilling bones or other animal matter. It is very impure, containing only about 10 per cent. of carbon.

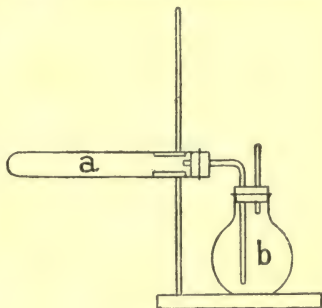


FIG. 65.

Properties of Amorphous Carbon.—Wood charcoal is porous—*i.e.*, full of tiny holes containing air. Owing to this enclosed air, charcoal floats on water, although its density is about 1.5 to 1.8.

EXPT. 162.—To show that Carbon is Denser than Water.—1. Powder some wood charcoal finely and shake with water in a test-tube. Allow it to stand, and observe whether it sinks or not.

2. Place a piece of charcoal in a boiling-tube with water, and while holding the charcoal below the water with a smaller tube, boil for a short time. Allow it to stand for a time; remove the smaller tube, and note what occurs.

3. Place a piece of charcoal in a beaker of water under the receiver of an air-pump. Exhaust the receiver and observe the charcoal.

Owing to its porosity, charcoal **absorbs gases**, and is used as a **disinfectant**.

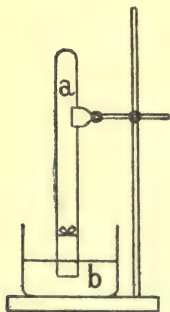


FIG. 66.

EXPT. 163.—To show that Charcoal absorbs Gases.—Fill a tube with dry ammonia-gas and invert it over mercury. Warm some fragments of charcoal (Why?) and introduce into the tube. Observe the rise of the mercury. Coconut-shell charcoal shows this best.

Another important property best shown by animal charcoal is its **bleaching** or **decolourizing** action.

EXPT. 164.—To bleach a Liquid with Charcoal.—Warm a little powdered boneblack with diluted claret, indigo solution or brown-sugar solution and filter it. The filtrate will be colourless. This is made use of

on the large scale in the manufacture of sugar and other substances.

If any of these impure forms of amorphous carbon are purified (sugar charcoal is the most pure) and then burned in oxygen, it yields **carbonic acid gas** and **no other product**. On account of its "liking" for oxygen, carbon is, as we learned in Expt. 82, used for reducing the oxide of metals to the metallic state. It is, indeed, the most widely-used commercial **reducing agent** (*coal*).

Crystalline Carbon.

1. **Graphite** (plumbago or blacklead) occurs as a **soft, black, lustrous, greasy** solid, usually crystallizing in hexagonal plates. It is used as a lubricant, polish, for "blacklead" pencils, and in the making of crucibles. Its **density is 2.5**, and, like other forms of carbon, burns to form **carbon dioxide only**. It is a good conductor of electricity.

2. **The diamond** is found in India, Borneo, South Africa, Brazil, etc., and is a **transparent, crystalline** solid. It is a **very hard, colourless** crystal; hence diamonds are used as gems, for cutting, grinding, and polishing other hard substances, for drills, and so on. The crystals are cubic or octahedral, like an alum crystal (*q.v.*), with a **density of 3.3 to 3.5**, and are non-conducting. Diamonds are infusible, but yield a kind of coke when heated away from the air, burning away completely in oxygen, to form **carbon dioxide only**.

Thus we see that carbon is an **element which possesses the property of existing in several distinct physical forms**.

This phenomenon is called **allotropy** from *allos*, other, and *trope*, change (*i.e.*, of condition). These separate and distinct physical substances, differing in density, hardness, and other physical properties, can be shown to consist of the same chemical element, carbon.

EXPT. 165.—**To show the Identity of the Forms of Carbon.**—Expt. 129 is repeated with a weighed amount of diamond, pure charcoal, or pure graphite. Such experiments always show that *one part of the different forms of carbon always yields 3.67 parts of carbon dioxide, and no other product—i.e., they must consist of the same chemical element, carbon.*

CHAPTER XVI

ACIDS—BASES—SALTS—NEUTRALIZATION

IN the preceding sections we have already learned something of the substances termed “acids”—*e.g.*, in Expt. 57 we found that dilute acids reacted with metals, liberating hydrogen; while Expt. 125 taught us that they reacted with carbonates, effervescence occurred, and carbonic acid gas escaped. (Do all acids behave in the same way?)

EXPT. 166.—To study the Behaviour of Acids.—Use *dilute* solutions of sulphuric, nitric, hydrochloric, citric, tartaric, and acetic acids, etc. (1) Taste; (2) test them with red and blue litmus; (3) add granulated zinc, warm and test for hydrogen; (4) add a little sodium carbonate and test for carbonic acid gas.

In Expt. 74 we found that the oxides of the non-metals were acidic—*i.e.*, formed acids when dissolved in water. (Is the acidity due to oxygen?) Lavoisier erroneously thought so; hence the name “oxygen” = acid-producer.

EXPT. 167.—To show that Acidity is due to Hydrogen.—Half fill a dish with dilute hydrochloric acid and add an excess of zinc. Test it from time to time, and find out whether the acid property ceases when all hydrogen is expelled.

Acids are not the only substances which lose hydro-

gen when treated with metals, for in Expts. 66 and 67 we found that the *neutral* substance, water, acted in a similar way. Thus, acids as a rule are sour-tasting substances, which turn blue litmus red, and react with metals and carbonates; and we define them thus: **An acid is a substance containing hydrogen, which can be replaced by a metal "forming a salt."** This is necessary in order to distinguish acids from such substances as water, which do not form salts.

The opposite of acid is **base**, and in Expt. 74 we learned that metals formed **basic oxides**. Moreover, in studying water, we met the substances caustic soda and potash, which turned red litmus blue; these were termed **alkalies**.

EXPT. 168.—To study the Behaviour of Bases.—Examine caustic soda, caustic potash, copper oxide, lead oxide, calcium hydroxide, magnesium hydroxide, and find out—

1. Whether they are soluble or not.
2. What taste they have. *Use very weak solution.*
3. Their effect on red and blue litmus.
4. Their effect on dilute sulphuric acid. Evaporate the neutral solution and examine the products.
5. If soluble, add their solutions to copper sulphate solution and note the effects.

Thus, all bases neutralize acid to form salts. Some are soluble, but not all. We therefore define a **base as the oxide or hydroxide of a metal, which can neutralize acids, to form salts and water only**. It is necessary to add *water only*, for many other substances may be used to prepare salts (*e.g.*, carbonates—see Expt. 125) which are not bases.

If a base is soluble—*e.g.*, caustic soda, caustic potash—it has a special name, **alkali**. Thus, an alkali is a

soluble base, and is always the hydrate or hydroxide of a metal, so that all alkalies are bases, but all bases are not alkalies.

The alkalies have a caustic taste, turn red litmus blue, and precipitate insoluble hydroxides from salt solutions.

EXPT. 169.—**To examine the Behaviour of Salts.**—Test the following *salts*, sulphate of soda, calcium chloride, bluestone, lead nitrate, common salt, sodium hydrogen sulphate, etc., in the following way:

1. Are they soluble in water or not ?
2. What effect have they on red and blue litmus solutions ?
3. What effect have their solutions on granulated zinc ?
4. What effect, if any, have they on washing-soda ?
5. What effect, if any, have they on copper sulphate solution ?

We find that salts as a rule are crystalline, soluble solids, neutral to litmus, with no acid or basic properties. Expt. 61 shows that the salt zinc sulphate is made of two parts—the metal zinc and the acid sulphuric acid. So we define **salts as substances formed when some or all of the replaceable hydrogen of an acid is replaced by a metal.** If all the acidic hydrogen is replaced, a **neutral or normal salt** is formed; but if only **part** is replaced, *some acidic* properties remain, and the salt is an **acid salt**, such as sodium hydrogen sulphate above. When two salts combine together, we obtain a **double salt**—*e.g.*, the alums, etc.—and if a salt is combined with an excess of a base, a **basic salt** results.

PREPARATION OF SALTS.**EXPT. 170.—To prepare Soluble Salts.**

METHOD 1. By the Action of Acids on Metals—A.
ZINC SULPHATE [A].—Half fill a dish with dilute sulphuric acid and add an excess of zinc. When all action is over, filter and evaporate the clear liquid until a spot crystallizes on a glass rod. Cool and examine the crystals as in Expt. 169.

EXERCISE.—Prepare the sulphate of magnesium, iron, etc., and describe the products you obtain.

B. LEAD NITRATE [B].—Half fill a dish with dilute nitric acid and warm. Add granulated lead until in excess, and proceed as in *A*.

EXERCISE.—Prepare the nitrates of zinc, magnesium, iron, etc., and describe your products. Prepare the chlorides of zinc, magnesium, iron, lead, etc., and describe your products.

METHOD 2. By the Action of Acids and Alkalies—
A. COMMON SALT (SODIUM CHLORIDE) [C].—Half fill a dish with dilute hydrochloric acid and add, spot by spot, a solution of caustic soda. Stir after each addition until a spot will not affect either red or blue litmus. Filter this neutral solution, if necessary, evaporate to crystallizing-point and cool. Examine the solid as before.

EXERCISE.—Prepare potassium and ammonium chlorides in this way, and examine them.

B. NITRE (POTASSIUM NITRATE) [D].—Repeat *A*, using a dish half full of caustic potash, and neutralizing with dilute nitric acid.

EXERCISE.—Prepare the nitrates of sodium and ammonium as above; also prepare the sulphates of sodium, potassium, and ammonium, and describe them.

METHOD 3. By the Action of Acids on Insoluble Bases.

—**A. BLUESTONE (COPPER SULPHATE) [E].**—Warm some dilute sulphuric acid in a dish, and add small quantities of black copper oxide until an excess remains. (Why ?) Filter, and proceed as above.

EXERCISE.—Prepare magnesium, zinc, ferrous, ferric, mercuric sulphates, and describe them.

B. MAGNESIUM NITRATE [F].—Repeat *A*, using dilute nitric acid and magnesium hydroxide.

EXERCISE.—Prepare the nitrates of copper, zinc, lead, calcium, etc., as above; also the chlorides of magnesium, lead, copper, zinc, iron, etc., and note their characteristics.

METHOD 4. By the Action of Acids on Carbonates—

CALCIUM CHLORIDE [G].—Half fill a dish with dilute hydrochloric acid and add chalk until no more effervescence occurs and a small excess of solid remains. Boil and filter, and proceed to obtain crystals as before.

EXERCISE.—Prepare as above the chlorides of copper, lead, zinc, magnesium, etc.; the sulphates of copper, zinc, magnesium, etc.; also the nitrates of copper, zinc, magnesium, lead, calcium, etc., and examine them.

EXPT. 171.—To prepare the Normal and Acid Sulphates of Sodium.—Take 25 c.cms. of dilute sulphuric acid and neutralize it with caustic soda solution [**H**], as in Method 2 (*A*) from a graduated cylinder, and note the volume of soda solution used—*e.g.*, 30 c.cms. Now take another 25 c.cms. of acid and add to it *half* the soda solution used before—*i.e.*, 15 c.cms. Evaporate these solutions to a *small bulk* and cool. When crystals have formed, dry them on filter-paper and test them (1) with litmus, (2) with zinc, (3) with washing-soda.

Acid Salts can also be prepared by the action of the *acid* on the *normal salt* [see Expts. 133 and 141 (2)]; so that we may proceed thus:

Place 5 gms. of strong sulphuric acid (3.5 c.cms.), or 25 c.cms. of dilute sulphuric acid supplied into a dish containing 16 gms. of sodium sulphate crystals dissolved in water. Evaporate the mixture to a small bulk, cool and crystallize, and examine as above, verifying the production of the acid sulphate of sodium.

Try the above experiments with hydrochloric, nitric, tartaric, phosphoric acid, etc. (Do all acids yield acid salts?) You will find that some acids, like hydrochloric and nitric acids, only yield **one salt**—*i.e.*, the normal or neutral salt—when treated with caustic soda or potash. Such acids are termed **monobasic**, and obviously contain only one part of acid hydrogen; other acids—*e.g.*, sulphuric and tartaric acids—yield **two salts**—a normal and an acid salt. Such acids are termed **dibasic**, and must, of course, contain two parts of acidic hydrogen. Phosphoric acid is an example of a **tribasic** acid, and yields **three salts**, because it contains three parts of acidic hydrogen.

EXPT. 172.—To prepare the Three Phosphates of Sodium.—1. Take 5 gms. of phosphoric acid, dissolve in water and neutralize it with a solution of caustic soda. Note the volume of the solution required—say 60 c.cms. Evaporate the solution to a small bulk, cool, and crystallize. Examine the product as above.

2. Repeat Expt. 1, but use only half as much caustic soda solution.

3. Repeat Expt. 1, but use one and a half times as much caustic soda solution. Compare the crystals obtained in each case. You will thus have three

phosphates [I]—viz., (1) disodium hydrogen phosphate, (2) monosodium dihydrogen phosphate, (3) trisodium phosphate.

All trials made by using less or more caustic soda solution fail to prepare any other salt; this thus proves that phosphoric acid is tribasic. Thus, the **basicity of an acid indicates** (1) the **number of salts** which may be derived from the acid; (2) the **amount of replaceable hydrogen present**; and also (3) the **amount of a base or alkali required to neutralize it**.

EXPT. 173.—To examine the Behaviour of Different Bases with Acid.—Place 25 c.cms. of dilute sulphuric acid in a dish; add 5 gms. of copper oxide and warm, stirring until the solution is neutral. Filter through two filter-papers, as in Expt. 43; wash until it is free from copper sulphate [L]. (How will you know?) Dry and weigh, and find out how much has been dissolved. Now evaporate the clear filtrate and allow it to crystallize. Meanwhile, take another 25 c.cms. of acid and add half as much copper oxide as was dissolved in the first case; warm until dissolved, and crystallize. Compare your results. (Do you obtain two different salts, as you did with caustic soda in Expt. 171?) Repeat this experiment with other bases. From these experiments we find that some bases will only form one salt with sulphuric acid—i.e., will replace *two* parts of hydrogen—whilst others form two salts—i.e., will replace *one* or *two* parts of hydrogen. **Bases** which can replace **one** part of acidic hydrogen are termed **monoacidic** or **monacid bases**; those which replace **two** parts of acid hydrogen are termed **diacid bases**, and so on. The **acidity of a base is thus very like the basicity of an acid**. This subject will be referred to in more detail later.

When the acidic properties of an acid are destroyed by the basic properties of a base, and *vice versa*, a neutral solution resulting from the process is termed **neutralization**. In all the above cases water was present throughout, and in order to show the formation of water during neutralization we must use an acid and a base which do not contain water.

EXPT. 174.—To show the Formation of Water during Neutralization.—Fit up the apparatus shown (Fig. 67); *a* is a hard tube containing a boat with dry

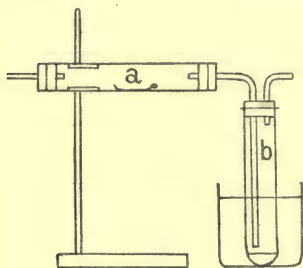


FIG. 67.

caustic soda, and fitted so that pure dry hydrochloric acid gas can be passed over this. Any moisture [*C*] that is formed collects in the receiver (*c*), kept cold by a beaker of cold water or ice.

Repeat the experiment with lead oxide, magnesium hydroxide, and other bases.

When a metal is placed in a dilute acid, the metal disappears (dissolves), and hydrogen escapes. This is an illustration of a type of chemical action called **replacement, or substitution**, since the metal takes the place of the hydrogen.

EXPT. 175.—To show the Replacement of Copper by Iron.—Clean a strip of iron, immerse it in a solution of copper sulphate and examine the iron. Repeat this, using a fair quantity of iron filings, and observe the change in colour of the solution. When all the blue colour has disappeared, boil the mixture to a small bulk; filter while hot, and allow to cool. Examine the crystals. (What are they?)

Here the iron has taken the place of the copper in copper sulphate, expelling the copper and forming iron sulphate [J]. (Can you suggest why this occurs?)

EXPT. 176.—To show the Replacement of Lead by Zinc.—Suspend a piece of zinc-foil in a solution of lead acetate and leave it for some time, examining it at intervals. A lead-tree is formed, owing to the zinc replacing the lead [K].

Similarly, when **two metals** exchange partners in **two salts**, and produce **two new salts**, the two original salts being decomposed, we term it **double decomposition**. This was used in Expt. 143 to prepare insoluble carbonates, and is used to **prepare all insoluble salts**. Numerous examples will be met with subsequently.

EXPT. 177.—To prepare Lead Chloride by Double Decomposition [L].—Add to a solution of lead nitrate some solution of a soluble chloride—*e.g.*, potassium chloride—spot by spot until no more precipitate is formed. Filter and evaporate the filtrate to a small bulk and obtain crystals. (Do these resemble the crystals of nitre you prepared above?) Wash the residue on the filter-paper; it is lead chloride. Thus, lead has changed its partner from nitrate to chloride, whilst the potassium changes over from chloride to nitrate.

EXERCISES.—Prepare copper carbonate from copper sulphate and washing-soda and potassium carbonate; barium sulphate from barium chloride and sulphuric acid and sodium sulphate; silver chloride from silver nitrate and hydrochloric acid and ammonium chloride, etc.

PREPARATION OF ACIDS.

As we have already learned, non-metallic oxides are **acid anhydrides** (see Expt. 74), and dissolve in water to form acids. Thus, we may **prepare acids by dissolving their anhydrides in water.**

Carbon dioxide and water give carbonic acid.

Sulphur dioxide and water give sulphurous acid.

Phosphorus pentoxide and water give phosphoric acid, and so on.

Caution.—To prepare Phosphoric Acid [M] and Sodium Phosphate from Phosphorus.—Take a wide glass tube, fitted as shown in Fig. 35, and fit it with a cork and tube at one end. Into the tube place a boat containing phosphorus, and pass a stream of air or oxygen through the apparatus. Place a beaker of water under the end of the tube, as shown, so that the water just covers the open end. Gently warm the phosphorus; it burns, forming phosphoric anhydride, which dissolves in the water, and forms phosphoric acid. When the phosphorus has completely burned, filter the liquid in the beaker, if necessary, and divide it into two portions. Evaporate one portion to dryness and examine the phosphoric acid produced. Boil the second portion, then neutralize it with caustic soda solution, as in Expt. 170 (2), and obtain crystals of the salt—sodium phosphate [I].

EXERCISES.—Prepare specimens of sodium carbonate from carbon, and of sodium sulphate from sulphur.

In later chapters a second method of preparing acids will be found. This is by **distilling their salts with strong sulphuric acid.** Thus a **nitrate** distilled with strong sulphuric acid yields **nitric acid**, while a **chloride** yields **hydrochloric acid**, and so on.

On p. 39 the list of elements was given under two headings—viz., non-metals and metals. We proceed with the information gained to compare these two classes of elements. Everyone knows the physical properties usually associated with various metals in bulk. Are metals solid? Are they dense? Do metals burnish or polish easily? What are wires usually made of? What are wires used for? Are metals tough and hard? Do you know of metal sheets? etc. We see that we usually consider **metals** as **hard, dense solids** (mercury is an exception), capable of acquiring a high polish—*i.e.*, exhibiting **metallic lustre**. They are **tough, ductile, and malleable**, and used for **conducting electricity and heat**. **Non-metals**, on the contrary, are **not all solids**, but include gases and liquids. They are **not hard**, they are **less dense**, and they cannot be polished. They are **brittle**, and **poor conductors of electricity**.

A cursory examination will show that not all metals possess all the properties ascribed to them above; moreover, some non-metals possess “metallic properties”—*e.g.*, graphite (carbon) possesses a high lustre, and gas carbon is a good conductor of electricity.

A table of a few of the commoner elements, with their approximate densities, is appended:

Non-Metals.

Oxygen ..	0.00144	Chlorine ..	0.00319	Phosphorus ..	1.80
Nitrogen..	0.00126	Iodine ..	4.95	Sulphur ..	2.00
Carbon..	1.9 to 3.5

Metals.

Aluminium ..	2.7	(Steel ..	7.80)	Sodium ..	0.98
Calcium ..	1.6	Lead ..	11.40	Silver ..	10.50
Copper ..	8.9	Magnesium ..	1.70	Tin ..	7.30
Gold ..	19.3	Mercury ..	13.60	Zinc ..	7.10
Iron ..	7.5	Potassium ..	0.87

We find, however, that these two classes of elements also show great differences in chemical behaviour. In the preceding section we have discovered (1) that **metals replace the hydrogen of acids to form salts** (Expt. 57), **non-metals do not**; (2) that whilst **metallic oxides are basic, non-metallic oxides are acidic** (Expt. 74); (3) that when solutions containing non-metals and metals are electrolyzed (Expts. 56, 61), the **metals are always liberated at the cathode, whilst non-metals are set free at the anode**. We shall also learn later (4) (see Expt. 255) that, whilst **metallic chlorides dissolve in water without decomposition, non-metallic chlorides are decomposed, forming acids**; and (5) that whilst the **non-metals form stable hydrides, the metals rarely form these compounds, and when formed, they are very easily decomposed**.

PREPARATION OF METALS.

Metals have been prepared in the preceding experiments in the following ways:

1. By **heating the oxides of mercury, silver, gold, and platinum**; but no other metallic oxides behave in this way (Expt. 68).

2. By **reduction of the oxides** (a) in hydrogen (Expts. 59 and 81)—*e.g.*, copper, lead, iron, etc.; (b) in coal-gas (Expt. 83); (c) with carbon (Expt. 82)—*e.g.*, lead, copper, silver; (d) with carbon monoxide (Expt. 157)—*e.g.*, iron, etc.

3. By **replacement with other metals** (Expts. 175 and 176)—*e.g.*, lead, copper, silver, etc.

4. By **electrolysis** (Expt. 61)—*e.g.*, zinc, magnesium, iron, copper, etc.

SOME COMMON METALS.

SODIUM.—Symbol, Na. Atomic weight, 23. Density, 0.98.

This metal, which we have used frequently in the foregoing experiments, is a soft metal, easily cut with a knife, showing a pinkish, lustrous surface, which rapidly becomes tarnished and dull, owing to the ease with which the metal is oxidized, a film of sodium peroxide being produced. Sodium has a remarkably low density, and easily floats on water, at the same time rapidly decomposing it. (Expt. p. 83. What is produced?) On account of this action **sodium** is usually kept under oil, and **must on no account be handled**. In Expt. 74 we found that it readily melted, forming a silvery liquid like mercury, and burned in oxygen with a golden-yellow flame, producing whitish fumes of sodium peroxide.

Sodium is widely distributed and is found in large quantities. The chief minerals are sodium chloride (rock salt) and sodium nitrate (Chili saltpetre), and it is generally manufactured by electrolyzing the former in the fused condition. Its most important compounds—the hydroxide, carbonates, chloride, nitrate, and sulphates—have been discussed, and these sections should be revised.

CALCIUM.—Symbol, Ca. Atomic weight, 40.

This metal, which resembles sodium in many ways, has only been manufactured on a large scale within the last few years: it is obtained by electrolyzing fused calcium chloride. It is a fairly hard metal, which slowly decomposes water, and forms calcium hydroxide and hydrogen. (Verify this.) After exposure to the air for some time it becomes coated with a hard grey solid, which consists of a mixture of the oxide and nitride.

Calcium occurs in enormous quantities as the carbonate (limestone, etc., p. 137), as the sulphate (gypsum), and phosphate (apatite, etc.). Its chief compounds are the oxide (lime), hydroxide, carbonates, chloride, and sulphate, which should be examined.

LEAD.—Symbol, Pb. Atomic weight, 207.

Lead is found in nature chiefly as the sulphide (galena) and also as the carbonate (cerussite). It is manufactured by roasting the former, lead and sulphur dioxide being pro-

duced. In the laboratory it is easily obtained by reducing its oxides in hydrogen, coal-gas, or with carbon (see Expts. 59, 82, 83) and by replacement (Expt. 176).

This well-known metal is a soft, grey, dense solid, which rapidly tarnishes in the air. It is very malleable, but not tenacious, and it owes its usefulness to the former property, to the fact that it is fusible (its melting-point is 330° C.), and also to its inactivity with water and acids (Expt. 57); nitric acid alone readily acts upon it. Lead forms several important alloys—*e.g.*, pewter, solder, etc. Its chief compounds are the oxides, nitrate, chloride, and carbonate, which should be studied.

IRON.—Symbol, Fe. Atomic weight, 56.

Iron is the most important metal to man, and is put to a large variety of useful purposes. The chief sources of this metal in nature are its oxides, hæmatite and magnetite, and the carbonate, spathic iron ore, clayband ironstone, etc. It also occurs in large quantities as the sulphide, iron pyrites. The metal is obtained by smelting together in a blast furnace hæmatite, limestone, and coal. In the laboratory it has been obtained by reducing the oxide in hydrogen, coal-gas, and carbon monoxide (Expts. 59, 83, 175).

Pure iron is too soft to be of use, but as commonly met with it contains small quantities of carbon, which gives it its peculiar characteristic properties. This impurity gives rise to the unpleasant-smelling gases accompanying hydrogen prepared from iron and dilute acids (Expt. 57). Iron rusts readily in damp air (Expt. 94), decomposes steam at a high temperature (Expt. 67), and dissolves readily in acids (Expt. 57). Its chief compounds, the oxides, ferric hydroxide, ferrous sulphate, ferric chloride, ferrous sulphide, and ferric sulphides, should be investigated.

MERCURY.—Symbol, Hg. Atomic weight, 200.

This element is the one metal, liquid at ordinary temperatures. On account of its high density it is used in barometers, and owing to its inactivity it is used for the collection of gases. It does not tarnish in air, but slowly combines with oxygen at high temperatures (p. 120). Dilute acids, save nitric, have no action upon it, although it dissolves in strong hot sulphuric acid. It unites readily with chlorine, sulphur, etc.

Mercury occurs in nature chiefly as the sulphide, cinnabar, from which it is easily obtained by roasting. Its chief compounds—the oxide, chloride, and nitrate—should be prepared and examined.

COPPER.—Symbol, Cu. Atomic weight, 63.

This red metal occurs in nature free, as its oxides and carbonates, but chiefly as the sulphide, associated with iron sulphide, as copper pyrites. Most of the copper is obtained by the Welsh process, in which a suitable selection of ores is roasted to produce pure cuprous sulphide. This is then roasted, and copper and sulphur dioxide are formed.

It is obtained in the laboratory by reducing its oxide in hydrogen or coal-gas (Expts. 59, 82), or by replacement (Expt. 175).

Copper is one of the most ductile and malleable of metals, and is much used in making alloys—*e.g.*, Dutch metal, brass, gun-metal, bell-metal, bronze, German silver, etc. It is an excellent conductor of electricity and heat, does not tarnish in dry air, and only very slowly in damp air. It is unaffected by dilute acids except nitric; it is, however, soluble in strong acids.

Its chief compounds are its oxides, chlorides, nitrate, and sulphate. These have been prepared and examined (pp. 180, 194), and should be revised here.

MAGNESIUM.—Symbol, Mg. Atomic weight, 24.

The chief minerals are the chloride (carnallite), the carbonate magnesite, and various silicates. It is mainly obtained from carnallite by the electrolysis of the fused magnesium chloride.

Magnesium is a silvery-white metal, having a density of 1.8, which burns brilliantly in air or oxygen. It decomposes steam (Expt. 66), reduces carbon dioxide (Expt. 128), and is soluble in acids (Expt. 57). Its chief compounds are the oxide and sulphate, which should be revised (pp. 180, 194).

ZINC.—Symbol, Zn. Atomic weight, 65.

This element occurs native as blende (the sulphide) and calamine (the carbonate). It is obtained from these by roasting, and then reducing the resulting zinc oxide by distillation with carbon (coal).

Zinc is a bluish-white, hard, crystalline, brittle metal. As usually met with it is *granulated* by melting and pouring into cold water. It is not easily oxidized, and is therefore used to make *galvanized* iron by dipping iron sheets in molten zinc. When heated strongly, it burns, forming zinc oxide, and is soluble in dilute acids and alkalies (Expts. 57, 85). Its chief compounds are the oxide, sulphate, and carbonate, which should be prepared and studied.

TIN.—Symbol, Sn. Atomic weight, 118.

Tin is a hard white metal, with a very low melting-point (232° C.), and is used to add fusibility to alloys—*e.g.*, solder. It is malleable, but not ductile. On account of the fact that it does not readily oxidize in air, it is used to coat iron in manufacturing tinware. It burns when strongly heated, is insoluble in dilute acids, but soluble in hot concentrated hydrochloric acid. With nitric acid it forms *tin oxide*, not the nitrate.

This metal occurs mainly as the oxide, tinstone, or cassiterite, from which it is obtained by reduction with coal. Its chief compounds are the oxides, chlorides, and sulphides, which should be examined.

PREPARATION OF BASES—*i.e.*, metallic oxides and hydroxides.

1. **Metallic Oxides.**—These are produced [see Expts. 21, 81, 91] by heating the metals in air or oxygen.

EXERCISES.—(1) Prepare small quantities of the bases formed from copper, lead, magnesium. (2) Prepare larger quantities of the bases formed from zinc, copper, iron, etc.

Basic oxides are also produced by heating the carbonates, hydrates, or nitrates of the metals (except sodium and potassium) (see Expt. 124).

EXERCISES.—(1) Prepare the bases zinc oxide, lime, litharge, etc., from carbonates. What escapes? (2) Prepare iron oxide, copper oxide, magnesium oxide, etc., from their hydrates. What escapes here? (3) Obtain the bases, lead oxide, copper oxide, calcium oxide, etc., from their nitrates. What is driven off in these experiments? Mention any other experiments in which bases are formed.

SOME COMMON METALLIC OXIDES.

SODIUM DIOXIDE is a yellowish basic peroxide, and is made by burning sodium in dry air or oxygen. It is soluble in cold water, and the solution, on warming, liberates oxygen and forms caustic soda, and is a vigorous oxidizing agent.

SODIUM MONOXIDE is a basic oxide, dissolving in water to form caustic soda. It is *said* to be produced by heating sodium in dry air at a low temperature.

CALCIUM OXIDE has already been investigated (Expt. 123). It is much used in its hydrated form (see below) in the preparation of mortar.

LEAD MONOXIDE (litharge) is made by heating lead in air, or by heating the carbonate or nitrate. It is a dense yellowish-red solid, basic, and slightly soluble in water, giving an alkaline reaction. It is used in the manufacture of flint-glass and is easily reduced to the metal, as in Expt. 59, by hydrogen, coal-gas, or carbon.

LEAD PEROXIDE (puce oxide of lead) may be obtained by the oxidation of litharge, or by the action of dilute nitric acid on red lead (see below). It is a basic peroxide, (and hence) on heating, liberates oxygen, and forms lead monoxide. When warmed with hydrochloric acid, it liberates chlorine, and with strong sulphuric acid oxygen is set free.

RED LEAD (minium) is made by heating litharge in air below 300° C. for forty-eight hours. It is used as a pigment, by glass-makers, plumbers, etc. When heated, it liberates a little oxygen, and forms lead oxide (Expt. 69). When warmed with hydrochloric acid and sulphuric acid, it gives chlorine and oxygen respectively—*i.e.*, it is a peroxide.

Expt.—Warm a few grams of red lead with dilute nitric acid, and collect the brown powder on a filter-paper. It is lead peroxide. Evaporate the solution, and crystallize; crystals of lead nitrate remain. Red lead is a compound or mixture of lead monoxide and lead peroxide.

ZINC OXIDE may be prepared by burning zinc in oxygen, or by igniting the hydrate, carbonate, or nitrate. It is a white, insoluble, basic solid, which turns yellow when heated, becoming white again when cooled.

MAGNESIUM OXIDE is a white, bulky powder, basic and insoluble. It may be prepared as in the case of zinc oxide.

CUPRIC OXIDE (black oxide of copper) is a black, basic, insoluble solid, obtained by heating the metal in air or oxygen, the carbonate or nitrate. It is also obtained by boiling a precipitate of the hydroxide in excess of alkali, and can be readily reduced to the metallic state by heating in hydrogen or coal-gas.

CUPROUS OXIDE (red oxide of copper) is a reddish solid, basic and insoluble. It is obtained by the reduction of cupric oxide or hydroxide; thus:

Expt.—Dissolve a little copper sulphate in water; add grape-sugar and excess of potash, and boil. A red precipitate of cuprous oxide is obtained. This substance dissolves in strong hydrochloric acid to form cuprous chloride, but with sulphuric acid it yields cupric sulphate and metallic copper.

TIN MONOXIDE (STANNOUS OXIDE) is a black solid, which, when heated in the air, readily forms.

TIN DIOXIDE (STANNIC OXIDE).—This may be obtained by heating tin in the air, or by the action of nitric acid on tin, and subsequent ignition of the hydrate formed.

FERRIC OXIDE may be obtained by heating the hydrate, nitrate, or sulphate. It is a red-brown basic solid, insoluble in water. Known as "rouge," it is used as a pigment and polishing substance. It may readily be reduced in a stream of hydrogen, coal-gas, or carbon monoxide.

MAGNETIC OXIDE OF IRON is a black solid, obtained by the action of steam on iron, or by burning iron in oxygen, and may be regarded as a combination of ferric oxide with a third oxide—

FERROUS OXIDE.—This is a basic oxide, very unfamiliar on account of the rapidity with which it is oxidized by the air.

MERCURIC OXIDE.—This substance is best obtained by cautiously heating nitrate of mercury, as a bright *red* powder, or by precipitation with potash, etc., as a *yellow* solid. It is a basic, insoluble solid, which is decomposed by heat (Expt. 68).

MERCUROUS OXIDE is obtained as a dark powder by adding caustic potash to a solution of mercurous nitrate. It changes on exposure into mercuric oxide and mercury.

2. **Metallic Hydroxides**—(A) **By the Action of the Metal on Water.**—We found that when the metals **sodium** and **potassium** acted on water, a caustic liquid was left, which, on evaporation, left the alkalies **sodium hydroxide** and **potassium hydroxide**.

Magnesium and iron behave differently. (What was formed with these ?)

Caustic soda (sodium hydrate or sodium hydroxide) and caustic potash may be prepared from the metals and water, as above, or from the oxides and water, but are more conveniently prepared thus:

EXPT. 178.—**To prepare Caustic Potash.**—Repeat Expt. 142, using potassium carbonate and one-quarter of its weight of lime. A white caustic solid is obtained. So that potassium carbonate and calcium hydrate yield potassium hydrate and calcium carbonate.

EXPT. 179.—**To examine the Caustic Alkalies.**—These are **white solids, strongly caustic, alkalies**. Expose a small piece of each to the air on watch-glasses and observe the moisture collected—*i.e.*, they are **deliquescent**. Place a piece of each in a c.cm. of water in a test-tube and notice how *hot* the liquid becomes as the solid dissolves. Rub a little on the finger; notice how **greasy** it feels. As we have already learned, these substances very rapidly **absorb** both **water** and **carbon dioxide** and other acid gases, and are therefore **used** (1) for **drying gases**, (2) for **removing and estimating carbon dioxide**, etc., from a gaseous mixture (see Expt. 132). They **neutralized all acids**, forming sodium and potassium salts (see Expt. 170).

(B) **By the Union of Water and the Oxide.**—In Expt. 118 we added water to quicklime and obtained the dry white solid slaked lime, or calcium hydroxide

Prepare magnesium oxide by slaking in this way. (C) Since **most metallic hydroxides are insoluble in water**, they are prepared by **double decomposition**.

EXPT. 180.—To prepare Copper Hydrate by Double Decomposition.—To a cold solution of copper sulphate add caustic soda until, on stirring, the clear liquid is alkaline to litmus; observe that it is colourless. (Why?) Filter off the precipitate and wash with cold water until the filtrate no longer affects red litmus. Allow the precipitate to dry; it is blue copper hydrate [N]. (What does the filtrate contain?) Verify your answer.

EXPT. 181.—To prepare Magnesium Hydrate.—Repeat the above experiment, using magnesium chloride and potassium hydrate, and boil before filtering. (Why?) Proceed as above. The residue is magnesium hydroxide [O]. (Why do we not boil in Expt. 180?) Try the effect of boiling. (What is formed?)

EXPT. 182.—To prepare Iron Hydrates.—Repeat Expt. 181, using the chlorides of iron (ferrous and ferric) and ammonium hydrate. Compare the precipitates of ferrous and ferric hydrates obtained.

SOME COMMON HYDROXIDES.

SODIUM HYDROXIDE (CAUSTIC SODA) has already been investigated (Expts. 178 and 179). It is prepared on the large scale by electrolyzing a solution of brine, and is largely used by soap and paper manufacturers. It liberates ammonia from its salts, and precipitates the hydroxides of many metals (Expt. 180).

POTASSIUM HYDROXIDE (CAUSTIC POTASH) closely resembles the sodium compound, and is prepared by exactly similar methods.

CALCIUM HYDROXIDE (SLAKED LIME) has already been studied (see Expt. 118). Its chief use is in the preparation of mortar—a mixture of slaked lime, sand, and water. The mixture hardens by drying and by the gradual absorption of carbon dioxide, to form calcium carbonate. (Verify this by testing some old mortar with dilute acids.)

MAGNESIUM HYDROXIDE is a white, basic, slightly soluble, (and alkaline) solid. It may be prepared by slaking the oxide, or by double decomposition.

ZINC HYDROXIDE is a white powder, obtained by precipitation with ammonia solution. It is soluble in caustic alkalies as well as in acids.

LEAD HYDROXIDE, obtained by double decomposition, is a white, slightly soluble solid, soluble in caustic alkalies.

FERROUS HYDRATE is precipitated as a dirty, greenish-white, solid when ammonia is added to a solution of a ferrous salt. It rapidly darkens, and ultimately becomes brown, owing to oxidation.

FERRIC HYDRATE is a reddish-brown solid, obtained by precipitation from a ferric salt. It is an insoluble, basic solid.

CUPRIC HYDROXIDE, prepared as in Expt. 180, is a pale blue, basic solid, which is quickly changed by boiling into the oxide.

Just as we prepared acid salts by adding acid to the normal salt, so we may prepare **double salts** by adding one salt to another.

EXPT. 183.—To prepare the Double Salt Potassium Cupric Sulphate.—Weigh out 15 gms. of potassium sulphate and 25 gms. of copper sulphate. Dissolve each in separate quantities, as little as possible, of boiling water; mix them and crystallize as usual. Examine the crystals and compare them with (1) potassium sulphate, (2) with copper sulphate.

EXPT. 184.—To prepare “Alum” from Clay.—Alum is a double salt formed by the union of potassium or ammonium sulphate with aluminium sulphate. Clay contains silica (sand) and alumina (oxide of aluminium)

—a base. We proceed thus: Heat 20 gms. of white china clay with 15 c.cms. of strong sulphuric acid just below the fuming-point, for some hours. Boil the product with about five times its weight of water; filter off the residue of silica and evaporate the solution to a syrup and allow it to cool. Crystals of aluminium sulphate are formed. Add 10 gms. of a strong potassium sulphate solution; warm until all is dissolved, and crystallize out the alum. Examine the product and try to grow a well-shaped crystal.

REVISION QUESTIONS.

CHAPTERS VII. AND VIII.

1. Name three substances of different types that give off oxygen when they are heated. Describe experiments to illustrate the chief chemical properties of oxygen.
2. The density of oxygen is to be determined from the following results: Weight of a hard-glass tube containing mercuric oxide = 28.736 gms.; weight of tube and contents after heating = 28.355 gms.; volume of oxygen = 283 c.cms. How would you carry out this experiment, and what is the density of the gas?
3. Of what use is "manganese dioxide" in oxygen mixture? Is it altered in any way? Give experiments to illustrate your answer.
4. How would you find the proportion of potassium chloride in a given sample of potassium chlorate? Give two methods, if you can, and say which is the most accurate, and why?
5. Of what use are reducing agents? Mention three common ones, and explain by an experiment the action of each.
6. How could you obtain some copper oxide from some metallic copper? and how would you regain the original copper?
7. How could you find the weight of oxygen in 1 lb. of (a) lead oxide, (b) mercuric oxide?
8. What is the result of heating (a) lead, (b) iron, in air? How would you verify your answer by experiment?

CHAPTER IX.

1. What do you understand by the terms *analysis*, *synthesis*? Explain either (a) how the composition of water may be shown by analysis, or (b) how it may be shown by synthesis. [C.L.J., 1909.]
2. How would you proceed in order to obtain about 20 c.cms. of water from hydrogen gas? Can this process be reversed—i.e., obtain hydrogen from water? If so, describe how this could be done.
3. What is the composition of water by weight? Describe accurately, with a sketch of the apparatus used, how this result could be verified.
4. How would you proceed in order to find out whether a given colourless liquid was water or not?
5. What does the name “hydrogen” mean? Give experiments to justify its choice.
6. “We may obtain metallic oxides from water, and we may also obtain water from metallic oxides.” What does this statement mean? Is it *accurately* expressed? Give experiments to illustrate it.

CHAPTERS X. AND XI.

1. Why does iron rust in moist air? Give experiments to illustrate your answer.
2. What are the results obtained when metals are heated in air? Give experiments illustrating your statements.
3. What happens when a candle burns in a closed bottle of air? Explain why the changes you state occur.
4. What do you know of combustion and respiration? Give experiments to illustrate your answer.
5. What are the four most important constituents of the air? Give reasons for your answer, and say how you could show by experiment the presence of the four you select.
6. “The composition of the air remains practically constant.” Why is this?
7. What important results followed Lavoisier’s experiments? Briefly describe how his experiments were carried out.
8. How would you determine the volume of oxygen and nitrogen present in the air if you were supplied with a bag of nails?

CHAPTERS XII. AND XIII.

1. "Matter is indestructible." Explain exactly what is meant by this statement, and describe how you would show that this is true (a) when a candle burns in air; (b) when iron dissolves in dilute sulphuric acid; (c) when limestone is treated with dilute hydrochloric acid ?
2. How could you obtain some quicklime from chalk, and how would you prepare from it some slaked lime and lime-water ? How could you reconvert lime-water into slaked lime and quicklime ?
3. What changes occur when a current of air is passed through a tube containing red-hot carbon and the issuing gas, passed through lime-water ?
4. Describe in detail how you would prepare carbon dioxide from chalk, and how you would show that the gas contains carbon and oxygen.
5. What is the composition of (a) quicklime, (b) slaked lime ? What chemical changes take place when slaked lime is boiled with dilute aqueous solution of sodium carbonate ?
[C.L.J., 1909.]
6. What is meant by "hard" and "soft" water ? Describe what happens when you breathe into solutions of (a) caustic soda, (b) slaked lime.
[C.L.J., 1907.]
7. How would you show that carbon dioxide is present in the air ? Is this connected in any way with the properties of natural waters ?
8. Describe exactly how you could find the volume of carbon dioxide in 1 kilogramme of any carbonate.
9. How would you show that equal weights of the two carbonates of sodium contain different quantities of carbon dioxide ?
10. What reactions take place when the following are heated in air—(a) chalk; (b) lead oxide (litharge); (c) lead peroxide; (d) mercury ?
[C.L.J., 1911.]

CHAPTERS XIV. AND XV.

1. What is meant by "reduction" ? Illustrate your answer by reference to experiments you have performed with hydrogen, carbon, and carbon monoxide. [C.L.J., 1909.]

2. How would you prepare (a) carbon monoxide, (b) carbon dioxide, from charcoal and air? Describe *exactly* how you would conduct the experiment in each case.
3. How would you prepare carbon monoxide? How would you distinguish experimentally between carbon monoxide, hydrogen, and oxygen? [C.L.J., 1907.]
4. Diamond, graphite (blacklead), are said to consist of the same element. Describe experiments to verify this. What is this phenomenon called?
5. Describe the destructive distillation of coal. Give experiments to show the chief properties of the gas produced.
6. Coal-gas is a good reducing agent. How would you verify this statement? What substances are produced in the experiment you describe, and what does it tell us of the composition of coal-gas?
7. How would you distinguish a piece of carbon from a lump of lead?
8. How would you distinguish a piece of diamond from a piece of glass?

CHAPTER XVI.

1. Explain the terms *acid* and *base*. Illustrate your answer by reference to sulphur dioxide, carbon dioxide, calcium hydroxide, and lead oxide. [C.L.J., 1909.]
2. Define "base," "acid," "alkali," "salt," and mention two examples of each. Describe the preparation of one member of each class.
3. Describe in detail how you obtain (a) copper oxide, (b) copper, from a solution of copper sulphate. [C.L.J., 1907.]
4. Define "base." Mention one member of each class, and describe its preparation and the result of its action with hydrochloric acid, nitric acid, and sulphuric acid.
5. What "alkalies" do you know? Describe their properties and their reaction with (a) nitric acid, (b) copper sulphate.
6. How would you test a given solid, to discover whether it was an acid or not?
7. What is a "salt"? Describe the chief methods of making them in the laboratory, and describe the products obtained in the experiments you quote.

REVISION TESTS.

F.

1. When a mixture of chlorate of potash with about one-fifth of its weight of black oxide of manganese is heated, the chlorate is converted into chloride, but the oxide of manganese remains practically unaltered. Give a full account of any experiments by means of which you would test the correctness of these statements. [C.W.B.S., 1911.]
2. Describe any means of showing that steam is decomposed by burning magnesium, and that the gas which is set free is the same as that evolved by the action of sodium upon water. Compare the appearance and properties of the other substances formed in the above reactions. [C.W.B.J., 1908.]
3. By what tests would you ascertain whether a given sample of powdered chalk contained any quicklime? Explain carefully how you could find the weight of chalk in a Kilogramme of such a mixture. [C.W.B.J., 1909.]
4. What is a "salt"? Describe fully the chief methods by which salts are produced in the laboratory, giving one example of each method. [C.W.B.J., 1909]

G.

1. How can a small quantity of oxygen in a pure state be obtained from atmospheric air? Describe the most accurate method you know of estimating the proportion by volume of oxygen in ordinary air. [C.W.B.J., 1908.]
2. Describe how you would convert copper into its black oxide, and how you would reduce this oxide again to metallic copper. Explain briefly how the latter process may be used to give the percentage composition of copper oxide and water. [C.W.B.J., 1909.]
3. A mixture of powdered chalk and slaked lime is shaken up, with water. Describe exactly how you would ascertain (1) how much, if any, of the mixture dissolves in each c.cm. of the water; (2) whether any dissolved matter is chalk or lime, or both. [C.W.B.J., 1910.]
4. Describe exactly how you would prepare from sodium hydrate specimens of (1) normal sodium carbonate. (2) acid sodium sulphate, (3) neutral sodium sulphate, (4) acid sodium carbonate. [C.W.B.S., 1906.]

H.

1. Experiment shows that atmospheric air is dissolved in water to a slight extent, and that the air so dissolved contains more oxygen than ordinary air. How would you proceed to investigate the truth of these two statements ? [C.W.B.S., 1909.]
2. How would you prepare a jar full of carbon dioxide from some washing-soda ? What experiments would you make in order to prove that carbon dioxide is composed of carbon and oxygen, and that this gas is a component of marble and chalk. [C.W.B.J., 1906.]
3. Describe briefly how you would prepare specimens of copper oxide, iron oxide, and magnesium oxide, and say what happens if these are heated (1) strongly with powdered carbon, (2) with hydrochloric acid. [C.W.B.J., 1906.]
4. Draw a diagram of the apparatus you would employ to obtain a quantity of atmospheric nitrogen free from oxygen and carbon dioxide, and explain how you would do it. [L.M., 1912.]

I.

1. Describe experiments to show that hydrogen and oxygen combine in definite proportions by weight *or* volume. How would you find experimentally the proportions of copper in copper oxide ? [C.W.B.J., 1906.]
2. How would you prepare some caustic soda if the only materials available were chalk and washing-soda ? Give the changes that occur in the course of the operations you describe. [C.W.B.S., 1906.]
3. Give examples to illustrate three different ways of preparing metallic oxides, and explain in any one case how you could recover the metal from its oxide with as little loss as possible. [C.W.B.J., 1910.]
4. If you were given distilled water and any other materials you might require, how would you prepare specimens of (1) permanently hard water, (2) temporarily hard water ? How could you soften the water again in each case, and show that you had been successful ? [L.M., 1912.]

J.

1. How would you show by experiment that water is composed of two colourless gases? By what further experiments would you show that one of these gases is a good reducing agent and the other a strong oxidizer?
[C.W.B.J., 1907.]
2. Sketch a simple apparatus for ascertaining the proportions by weight in which carbon and oxygen unite to form carbon dioxide, and explain how you would carry out the experiment and deduce the result. [C.W.B.S., 1911.]
3. How would you show that (a) water is formed when dry hydrogen chloride is passed over dry lead oxide; (b) that carbon dioxide is formed when alcohol is burnt in a spirit-lamp; (c) that oxygen is liberated when lead nitrate is heated? Draw the apparatus you would use in each experiment. [L.M., 1910.]
4. Describe some experiments by which metals may be prepared in the free state from their oxides. Give neat diagrams of the apparatus you would employ.
[C.W.B.J., 1907.]

K.

1. How may quicklime be employed to detect the presence of two of the four most important constituents of the ordinary atmospheric air and metallic iron to prove the presence of the remaining two? Describe fully the actual experiments you would make. [L.M., 1910.]
2. What is meant by a reducing agent? Mention three such agents, and describe an illustrative experiment in each case. [C.W.B.S., 1909.]
3. How is sodium bicarbonate related to sodium carbonate? How would you prepare specimens of these salts from sodium hydroxide? Give some account of their properties, and explain for what purposes they are used. [L.M., 1911.]
4. What do you understand by the *basicity* of an acid? Illustrate your answer by reference to actual examples. If you were given some sodium carbonate and sulphuric acid, how would you prepare a specimen of crystallized sodium hydrogen sulphate?
[L.M., 1911.]

PRACTICAL EXERCISES.

1. Determine the weight of the substance *A* required to liberate 100 gms. of carbon dioxide when heated.
2. Determine the loss in weight when 100 gms. of *B* is heated in air.
3. Determine the percentage loss of weight when the substance *C* is heated.
4. Describe all that you observe when the substance *D* is heated. Examine and identify the products of the action.
5. The substance *E* is a mixture of two substances, one of which is soluble in water and the other is not. Separate and identify them. Apply as many confirmatory tests as you can. [C.L.J., 1911.]
6. Identify the compound *F*. [C.L.J., 1909.]
7. Find which of the following solvents will dissolve *G*: Water, dilute hydrochloric acid, dilute nitric acid, and dilute sulphuric acid. Describe carefully the experiments you make to find out whether the substance has dissolved or not. [C.L.J., 1911.]
8. Heat *H* in a dry test-tube, and identify the volatile products. What is the reaction of *H* towards litmus (*a*) before being heated, (*b*) after being heated? Make a solution of *H*. Make also a solution of the substance left after heating *H*. Pour the solutions together, and describe what you observe. [C.L.J., 1911.]
9. Examine and describe carefully what you observe when (*a*) caustic soda solution, (*b*) dilute ammonia, are slowly added to a dilute solution of *I*.
10. Describe the appearance of the substance *J*, and find its density. Heat a portion of it strongly for five minutes; cool and compare the action of water, litmus, and dilute acids on the original substance and the residue. State what you think has taken place.
11. Find what occurs when the metal *K* is placed in a solution of the solid *L*.
12. Find what occurs when the solution *M* is added to a solution of the solid *N*.

PART III.

CHAPTER XVII

SULPHURIC ACID AND ITS PRODUCTS

WE next proceed to examine the acids and their products. Perhaps the best known and most important is **sulphuric acid**, or **oil of vitriol**. It owes its common name to the original production of it by the action of heat on certain *glassy* crystals—notably **green vitriol**, or **ferrous sulphate**.

EXPT. 185.—**To study the Action of Heat on Green Vitriol.**—Place some powdered ferrous sulphate in a small ignition-tube, provided with a right-angled escape-tube (Fig. 56), and strongly heat. Observe that steamy fumes escape, which may be condensed in a cooled test-tube, whilst a reddish powder remains in *a*. Examine this residue; it is **ferric oxide** [**A**] (**rouge**). (How would you verify this statement?) Test the liquid obtained (1) with litmus; it is **acid**; (2) by adding dilute hydrochloric acid and barium chloride; a white precipitate is produced. The liquid obtained is **sulphuric acid**.

The further preparation of this acid is discussed later (see p. 216, etc.).

EXPT. 186.—**To study the Properties of Sulphuric Acid.**—It is an **oily, colourless liquid** (*oil of vitriol*),

and is **extremely dangerous to handle**, destroying flesh, clothes, paper, etc., with which it comes in contact, and the **greatest care must be taken when using it**.

1. Pour one spot of the strong acid into a test-tube full of water; taste it, and test it with blue litmus.

2. Place a few drops of the strong acid in a porcelain dish, and weigh. Expose it to the air for half an hour and reweigh. (What do you observe, and how do you explain it ?)

Liquids which, like sulphuric acid, are fond of water are said to be **hygroscopic**, and are very useful for **drying gases**.

3. Pour strong sulphuric acid into a little water in a test-tube, and note the **heat** developed.

4. Pour some strong sulphuric acid into a saturated solution of sugar [B], and observe the steam given off and the resulting porous mass of carbon.

5. Place a piece of wood or paper into some strong sulphuric acid, and observe the charring effect.

6. Place a crystal of copper sulphate [C] in some strong sulphuric acid, and observe that it loses its colour. (Why ?)

These experiments show how powerful is the **dehydrating—i.e., water-removing—action** of sulphuric acid.

7. Like all acids, it neutralizes alkalies and bases, and decomposing carbonates forming the **salts** known as **sulphates** (see Expts. 170 and 171).

8. Pour **dilute sulphuric acid** on fragments of magnesium, zinc, iron, etc., in test-tubes, and observe that **hydrogen is liberated, sulphate of the metals being formed**.

9. **Warm (but do not boil)** a small quantity of strong sulphuric acid with fragments of the above metals,

and test the gas evolved by smell, wet litmus-paper, and a light. No hydrogen is given off, but a pungent-smelling gas is noticed (like burning sulphur : it is **sulphur dioxide**), which is acid, and bleaches litmus.

Very few metals with strong sulphuric acid liberate hydrogen. Sodium will, but this experiment must be done by the teacher. Now, when strong sulphuric acid is strongly heated—*e.g.*, when it is dropped on to red-hot porcelain—it is found to break down into water, sulphur dioxide, and oxygen. Hence strong sulphuric acid may, and does, act as an *oxidizing agent*, and it is due to this fact that hydrogen is not liberated in the above action on metals.

10. Warm some **strong sulphuric acid** with various **chlorides**, and test the steamy fumes with litmus; they are acid, for **hydrochloric acid gas** and **sulphates** are formed.

11. Repeat Expt. 10 with **nitrates**, dropping in a little copper wire or foil. A brown gas is liberated, showing that **nitric acid** was formed; a **sulphate** is also produced. Similar results are obtained with other salts. Thus, **sulphuric acid** is used for preparing **hydrogen**, **carbon dioxide**, **sulphur dioxide**, **carbon monoxide**, etc., and **acids** in general.

EXPT. 187.—**To examine the Properties of Sulphates.**—Prepare as many sulphates as you can (Expts. 171 and 172), examine your products and verify—

1. **Most are soluble crystalline solids**; the sulphates of lead, barium, calcium, and strontium are insoluble.

2. The soluble salts when in solution **react with barium chloride** solution, producing a **white precipitate of barium sulphate** [D], insoluble in dilute hydrochloric acid.

3. Mix a sulphate with powdered charcoal and heat a little of the mixture strongly. Cool, and add a spot

or two of dilute acid to the product. Note that an offensive-smelling gas is evolved; it is hydrogen sulphide. This gas is liberated from sulphides (*q.v.*) by acids, and we may thus note that **sulphates are reduced by carbon to sulphides.**

SOME COMMON SULPHATES.

A. THE VITRIOLS.—1. **GREEN VITRIOL (FERROUS SULPHATE)** is prepared by the action of sulphuric acid on the metal (Expt. 170). It is a green, glassy crystal, soluble in water, and absorbs nitric oxide (*q.v.*), forming a brown solution. When strongly heated, sulphuric acid is produced (Expt. 185), whilst ferric oxide remains.

2. **BLUE VITRIOL (COPPER SULPHATE)** is obtained by the action of sulphuric acid on the metal oxide, hydrate, carbonate, or nitrate. It is a blue, well-crystallized, soluble solid, which, on cautiously heating, becomes white. (What does it lose?) This white residue is used as a test for moisture, by which the blue compound is re-formed.

3. **WHITE VITRIOL (ZINC SULPHATE)** is obtained from the metal, oxide, carbonate, etc., when treated with sulphuric acid. It is a very soluble, colourless, glassy, prismatic, crystalline solid.

B. FERRIC SULPHATE is obtained by dissolving ferric oxide or hydroxide in sulphuric acid, or by oxidizing ferrous sulphate with nitric acid. It is a pale-coloured, soluble, amorphous solid, which, with alkali sulphates, forms well-crystallized *iron alums*.

MAGNESIUM SULPHATE closely resembles zinc sulphate, and it is prepared like the latter. It is more commonly known as *Epsom salts*, and is used as a purgative.

SODIUM SULPHATE and **SODIUM HYDROGEN SULPHATE** were prepared and examined in Expt. 171.

POTASSIUM SULPHATE and **POTASSIUM HYDROGEN SULPHATE** are prepared like the corresponding sodium salts, and are very similar.

CALCIUM SULPHATE occurs in Nature as gypsum, selenite, and alabaster. It is a white, almost insoluble, solid, and is hence prepared by double decomposition (Expt. 177). To

the presence of this substance in solution natural waters mainly owe their permanent hardness. *Plaster of Paris* is obtained by heating gypsum for some time at 130°C . The dry solid is mixed with water to form a paste, and sets, owing to the rehydration of the calcium sulphate.

LEAD SULPHATE and BARIUM SULPHATE are white, insoluble solids, produced by double decomposition.

Sulphur Dioxide.

Chemical formula, SO_2 . Molecular weight, 64. Density, 32 (or 0.00288 gm. per c.c.).

This gas, as we have learned already, is produced when sulphur burns in air [E] or in oxygen. It is also obtained when sulphides are roasted, the oxides of the metal and sulphur both being formed. This is a method used on the large scale for preparing metallic oxides and sulphur dioxide.

EXPT. 188.—To examine the Effect of roasting Iron Pyrites.—Place some powdered iron pyrites in the hard-glass tube (Fig. 27), and when hot, pass oxygen over the substance. Observe that a gas with a pungent odour is set free; it is sulphur dioxide. Examine the residue in the tube; it is reddish-brown iron oxide [F]. (How would you prove this ?)

EXPT. 189.—To study the Action of Metals and Non-Metals on Strong Sulphuric Acid.—Repeat Expt. 186 (9), using mercury, sulphur, carbon [G], and test the gas by smelling and with wet blue litmus. In every case sulphur dioxide is formed, sulphates of the metals and oxides of the non-metals being also produced.

It is therefore purely a matter of convenience which we use.

EXPT. 190.—To prepare Sulphur Dioxide.—Copper is the metal generally employed. Place some copper

turnings or wire in the flask (a), fitted with a thistle funnel and delivery-tube. The latter passes into a two-necked bottle containing strong sulphuric acid (Why?), and the dry gas is collected by displacing the air upwards. (What property of the gas does this illustrate?) If

the gas is required free from air, it is collected over mercury. Strong sulphuric acid is poured down the thistle funnel, and the flask heated. A vigorous action occurs, and the gas is rapidly given off. When the action is over, cool the flask, pour in water, warm, and filter.

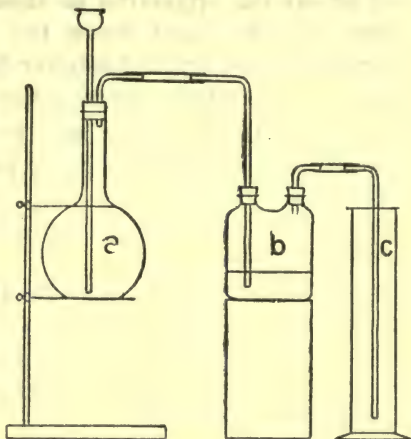


FIG. 68.

Evaporate the filtrate until crystallization-point is reached, and cool. Blue crystals [H] of copper sulphate will be found. The gas obtained in this way is not pure, but contains a little hydrogen, sulphide, etc.

Close examination of the residue in the flask will reveal the fact that a black substance remains. This, on examination, is found to be copper (cuprous) sulphide (How would you verify this?), produced by the reduction of the copper sulphate or sulphuric acid by the nascent hydrogen first formed.

EXPT. 191.—To study the Action of Dilute Acids on Sulphites.—Add to small quantities of various sulphites some dilute sulphuric and hydrochloric acids. Warm and test the gas. You learn that sulphites,

like carbonates, are decomposed by dilute acids liberating sulphur dioxide, and forming a salt and water. We may thus prepare sulphur dioxide from sulphites.

EXPT. 192.—**To prepare Sulphur Dioxide from Sulphites.**—Place some sulphite of soda in the flask (*a*), and fit up the apparatus as shown (Fig. 68). Pour dilute sulphuric acid down the thistle, and warm. Effervescence occurs, and sulphur dioxide[I] is given off. This is a convenient method if the gas is required pure.

EXPT. 193.—**To study the Properties of Sulphur Dioxide.**—Sulphur dioxide is a pungent-smelling gas, with an acid taste.

1. Place a wet litmus-paper in the gas; it is at once reddened—*i.e.*, is **acidic**.

2. It is, as we see by its collection, **denser than air**. Its density compared with air = 2.2. Find its density as in Expt. 59, and compare it with hydrogen. Its density is 0.00288 gm. per c.cm., or **thirty-two times as dense as hydrogen**.

3. Bubble the gas through weak magenta solution, or place an azalea or primrose in a jar of the gas. In both cases the colour is discharged—*i.e.*, the gas is a **bleaching agent**, and in general it **bleaches by reduction**.

4. Bubble the gas through a solution of potassium bichromate [J]; the orange colour is changed to green, the bichromate being *reduced* to chrome alum.

Repeat this experiment with potassium permanganate; the colour is completely discharged, owing to the **reducing action of sulphur dioxide**.

5. Bubble the gas through a weak solution of bleaching-powder, and observe that the odour of chlorine disappears. Sulphur dioxide is therefore used as an **antichlor**, to remove chlorine from a fabric after bleaching. Sulphur dioxide also destroys the

organisms of putrefaction and decay, and is on this account used as a **disinfectant** and **preservative**.

6. Pass dry sulphur dioxide through a U-tube surrounded by a freezing mixture of ice and salt. A colourless liquid collects, which has the odour of sulphur dioxide, and boils at -10°C . It is **liquid sulphur dioxide**.

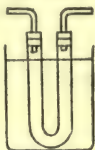


FIG. 69.

7. Bubble some sulphur dioxide through water. Test the liquid (1) by smell, (2) with litmus, (3) neutralize a little with caustic soda, and evaporate. A salt is formed, so that **sulphur dioxide dissolves in water to form an acid—viz., sulphurous acid [K]**. Hence the gas cannot be collected over water.

EXPT. 194.—To find the Composition by Weight of Sulphur Dioxide.—A boat (Fig. 59, *d*) containing sulphur is weighed and placed in the hard-glass tube containing a column of copper oxide (*c*). (Why?) Attached to it is a weighed U-tube of soda lime. *c* is made red-hot, and a current of oxygen passed through the apparatus. The sulphur is gently heated; it burns, forming sulphur dioxide [**A**], any sulphur passing over being oxidized by the copper oxide. The sulphur dioxide is retained in the U-tube. When complete, the apparatus is allowed to cool; *d* is reweighed, and then the U-tube.

RESULTS—

Weight of boat and sulphur before = 3.33 gms.

Weight of boat and sulphur after = 2.31 „

∴ Weight of sulphur burned = 1.02 gms.

Weight of U-tube after = 45.37 gms.

Weight of U-tube before = 43.32 „

∴ Weight of sulphur dioxide = 2.05 gms.

Thus, 2.05 gms. of sulphur dioxide contain 1.02 gms. of sulphur and 1.03 gms. of oxygen;
1 gm. of sulphur dioxide contains

$$\frac{1.02}{2.05} = 0.5 \text{ gm. of sulphur.}$$

and

$$\frac{1.03}{2.05} = 0.5 \text{ gm. of oxygen.}$$

or half the weight of sulphur dioxide is sulphur and half is oxygen.

***EXPT. 195.—To find the Composition of Sulphur Dioxide by Volume.**—Repeat Expt. 130, using a small piece of sulphur in a porcelain boat. It will be found that sulphur dioxide contains its own volume of oxygen. Since the relative densities of these gases are 32 and 16, it follows that—

1 volume of sulphur dioxide weighing 32 parts by weight contains 1 volume of oxygen weighing 16 parts by weight; and hence contains sulphur weighing 16 parts by weight.

So that sulphur dioxide consists of half its weight or 50 per cent. of oxygen and 50 per cent. of sulphur, thus confirming the result obtained in Expt. 193.

Sulphurous Acid forms the salts **sulphites**.

Compare (1) sulphur dioxide with carbon dioxide; (2) sulphurous acid with carbonic acid; and (3) sulphites with carbonates.

This acid, which is formed, as above, by dissolving sulphur dioxide in water, forms two classes of salts—**normal** and **acid salts**—and is therefore **dibasic**. These are prepared like the corresponding carbonates.

EXPT. 196.—To prepare Sodium Hydrogen Sulphite.—Pass sulphur dioxide through caustic soda solution

until the odour is quite marked [L]. Sodium acid sulphite results.

EXPT. 197.—To prepare Sodium Sulphite.—Add $2\frac{1}{2}$ parts of sodium hydrogen sulphite to 1 part of sodium hydrate, and evaporate; normal sodium sulphite [M] is formed.

Many sulphites are insoluble, and may be prepared by double decomposition—*e.g.*, on adding a solution of barium chloride to a solution of sodium sulphite a white, insoluble substance—barium sulphite [N]—is precipitated, while sodium chloride remains in solution.

EXPT. 198.—To examine the Properties of Sulphites.

—1. Test the solubility of as many sulphites as you can. Do you learn that sulphites are **insoluble in water**, except sodium and potassium? What method of preparation must you use for these?

2. Heat as many sulphites as you can in a small tube, testing the gas evolved. Do most sulphites give off sulphur dioxide and leave the oxide? Why is it that **only the acid sulphites give off sulphur dioxide when heated**, and leave sulphites?

3. Warm sulphites with dilute acids; sulphur dioxide is given off, a salt and water remaining.

4. Place fragments of sulphites into acidified permanganate or bichromate of potash solution; a **reduction occurs**. Thus, sulphites are reducing agents.

EXPT. 199.—To study the Effect of Oxidation on Sulphites and Sulphurous Acid.—Expose solutions of sodium sulphite and of sulphurous acid to the air for some time, and observe the gradual disappearance of the pungent smell. Test small portions with acid potassium permanganate solution, or bichromate. At last **no reduction occurs**. Test the remaining sul-

phurous acid liquid with litmus; it is still acid. Thus, oxidation has occurred, resulting in a new salt (**sodium sulphate**) and a new acid (**sulphuric acid**) [O].

This oxidation can be brought about by many oxidizing agents. Warming with nitric acid (*q.v.*), potassium bichromate, potassium permanganate, chlorine water, and bleaching-powder (*q.v.*), all effect this change.

EXPT. 200.—To compare Sulphates with Sulphites.—Make solutions of sodium sulphate and a *freshly cleaned* crystal of sodium sulphite (Why?); and test as below:

1. Warm each with a little dilute hydrochloric acid; the sulphite gives off sulphur dioxide, the sulphate does not.

2. Add to each a little acidified potassium bichromate or permanganate. Reduction occurs with the sulphite, but not with the sulphate.

3. Add to each some barium chloride solution. Each gives a white precipitate. (What are they?) Add some hydrochloric acid; the sulphite precipitate disappears (being decomposed), the sulphate precipitate is insoluble.

******Now that we have learned that sulphurous acid—*i.e.*, a solution of sulphur dioxide—in water can be oxidized readily and easily to sulphuric acid, the question arises, *Can sulphur dioxide gas be oxidized?*

EXPT. 201.—To oxidize Sulphur Dioxide.—Fit up the apparatus shown in Fig. 70. It consists of a hard-glass tube containing platinized asbestos at *b*. (Platinized asbestos is made by saturating asbestos

with platinic chloride, and strongly heating. It is really asbestos covered with very finely divided platinum.) Oxygen and sulphur dioxide, dried by bubbling through strong sulphuric acid, are passed over the gently heated platinized asbestos. Dense white fumes escape, which may be condensed in a well-cooled flask or tube, as shown, to white silky crystals; they are crystals of sulphur trioxide. Thus, sulphur dioxide has been oxidized to sulphur trioxide [P].

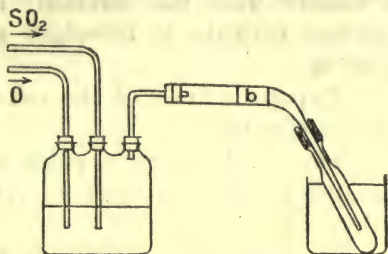


FIG. 70.

** Sulphur Trioxide.

Chemical formula, SO_3 . Molecular weight, 80 [= 32 + 3 (16)]. Density of vapour, 40.

This substance is usually prepared as above, but it may also be obtained by dehydrating sulphuric acid—i.e., by removing water from it—or by heating Nordhausen (fuming) sulphuric acid.

EXPT. 202.—To obtain Sulphur Trioxide from Sulphuric Acid.—Place in a small retort 2 parts of phosphorous pentoxide [Q], and add to it 1 part of strong sulphuric acid. After letting it stand for some time, gently distil the mixture, condensing the fumes as in Expt. 201 above.

EXPT. 203.—To examine the Properties of Sulphur Trioxide.—As found above, sulphur trioxide forms white silky crystals, which are very volatile, and form

dense fumes in air. (They melt at 15° C., and boil at 46° C.)

1. Drop a small portion into water; a hissing sound is heard. Test the solution; it is sulphuric acid. Sulphur trioxide is therefore **sulphuric acid anhydride [R]**.

2. Expose a little of the solid to the air; it **deliquesces** rapidly.

3. Place a little on a piece of paper, sugar, etc.; **charring occurs**, owing to its vigorously withdrawing water from organic substances.

N.B.—Care must be taken to keep it from the skin, or painful sores result.

4. Pass the vapour through a hard-glass tube containing some dry barium oxide, and note that they unite with incandescence **[S]**, forming barium sulphate.

5. Pass the vapour slowly through a red-hot tube; no dense fumes escape. **Test the escaping gases**, and verify the fact that sulphur trioxide is decomposed into sulphur dioxide and oxygen **[T]**.

Owing to the extreme readiness with which this substance unites with water, it is usually kept in sealed tubes or flasks.

****Sulphuric Acid.**

Chemical formula, H_2SO_4 . Molecular weight, 98 [$= 2(1) + 32 + 4(16)$].

As we can now prepare sulphur dioxide (Expts. 188-191) and oxidize it (Expts. 199 and 201), to form sulphuric anhydride, we are in a position to prepare this most important acid.

Sulphuric acid is being manufactured in ever increasing amount by what is known as the contact

process. Sulphur dioxide prepared by roasting pyrites (Expt. 188) and air are specially purified, to remove moisture and carbonic acid gas. The purified gases are then passed through pipes containing platinized asbestos. The resulting sulphur trioxide [P] is dissolved in water, and the sulphuric acid produced is concentrated by distillation.

An older process, still largely employed, is the so-called **lead chamber process**. In this sulphur dioxide, produced by roasting iron pyrites in the furnace [P], together with nitric acid, generated from nitre and

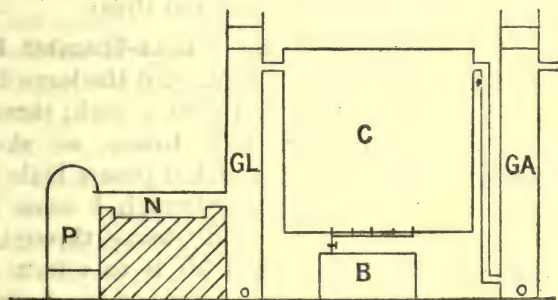


FIG. 71.

sulphuric acid in the nitre-pots (N), pass up the tower (Glover's tower) into the lead chamber (C). Here it meets with air and steam, injected from the boiler (B). Sulphuric acid is produced, and the escaping gases pass up the second tower (Gay Lussac's), where the nitrogen oxides meet a stream of sulphuric acid and are dissolved. This acid is pumped to the top of the Glover tower, and as it slowly trickles down, meeting the hot furnace gases, the nitrogen oxides are set free, and swept once more into the lead chamber.

The changes that occur are still not quite under-

stood, but are probably as follows [U]: The sulphur dioxide and nitric acid react, forming sulphuric acid and nitric oxide; the latter gas forms nitrogen peroxide with the air in the lead chamber. This nitrogen peroxide now oxidizes a further quantity of sulphur dioxide, which dissolves in water, forming sulphuric acid, and is itself reduced again to nitric oxide. The process is then repeated, so that a small quantity of nitric oxide converts a very large quantity of sulphur dioxide into sulphuric acid, which is then purified and distilled.

This manufacture may be imitated thus:

EXPT. 204.—To illustrate the “Lead-Chamber Process” for preparing Sulphuric Acid.—Fit the large flask

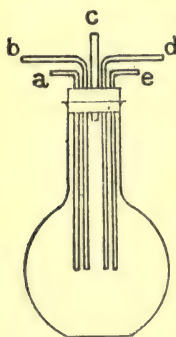


FIG. 72.

(15 to 20 litres) with a cork, through which pass five tubes, as shown (Fig. 72). Through *a* pass a little dry sulphur dioxide, through *b* some dry nitric oxide, and steam through *e*. The fourth tube (*d*) is to admit air. When the nitric oxide is first admitted, the flask becomes full of brown fumes. After a time these disappear, and if too much steam has not been added, crystals will be produced, known as “chamber crystals” [V] (nitro-sulphonic acid). These immediately dis-

appear when more steam is added, sulphuric acid being formed, and brown fumes once more fill the flask. It is extremely doubtful if chamber crystals are ever really produced on the large scale, as the supply of steam is too generous.

CHAPTER XVIII

SULPHUR

CHEMICAL symbol, S. Atomic weight, 32.

This element, more familiarly known as **brimstone**, occurs free in volcanic regions (Sicily, etc.), and combined with metals forms important **ores**—viz., sulphides (*e.g.*, galena, lead sulphide; blende, zinc sulphide; pyrites, iron sulphide; etc.). It also occurs in large amounts as sulphates—*e.g.*, heavy spar, barium sulphate; gypsum, calcium sulphate; etc.—and is present in small quantities in animal tissues and in plants—*e.g.*, garlic, horse-radish, and so on.

EXTRACTION OF SULPHUR.—The purest form of sulphur is that collected near volcanoes, but it is associated with earthy substances, usually limestone and gypsum, etc. It is separated from these by melting in pots or on a sloping hearth, when the sulphur melts and drains away from the impurities. This crude sulphur is then purified by distillation as below: The sulphur is melted in a pot (*A*) (Fig. 73), and runs into the retort (*R*), where it is boiled; the vapour condenses in the brickwork chamber (*C*). At first **flowers of sulphur** form on the walls, but as *C* becomes hotter, the whole melts and collects on the floor, from whence it is run off at *P* into moulds, in the form of **roll sulphur**, or **block sulphur**.

EXPT. 205.—To obtain Sulphur from Pyrites.—Place some iron pyrites in a small tube, closed at one end, and heat. Observe the globules of sulphur [A] on the cool sides.

On the large scale heaps of pyrites (2,000 tons or more) are heated, but much of the sulphur is lost in this process.

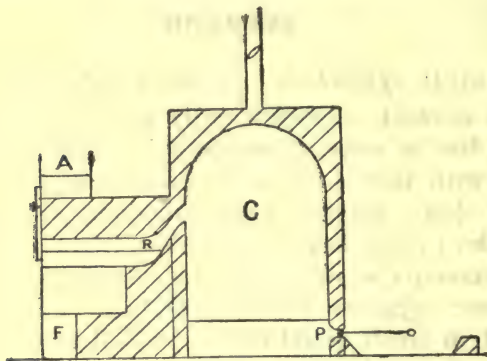


FIG. 73.

EXPT. 206.—To study the Properties of Sulphur.—Sulphur is a **lemon-yellow solid**, very **brittle**, and a **poor conductor** of heat and electricity.

1. Hold a piece of roll sulphur in the warm hand near the ear; a crackling noise is heard, due to the warm outer layer expanding and separating from the cold inner layer.

As we found in Expt. 26, sulphur is **insoluble in water**, but readily soluble in **carbon disulphide**. In its chemical properties sulphur is closely related to oxygen; it readily unites with **non-metals and metals forming sulphides**.

2. Repeat Expt. 74 (2). Sulphur **burns** readily

in air or oxygen with a blue flame, forming sulphur dioxide [B].

3. Bubble hydrogen through gently boiling sulphur (Fig. 74); an evil-smelling gas is produced—it is hydrogen sulphide [C], or sulphuretted hydrogen.

4. Boil some sulphur in a small flask or test-tube, and plunge into the vapour a roll of hot copper foil; the copper burns in sulphur vapour, forming copper sulphide [D]; iron readily unites to form iron sulphide [E].

Carbon and sulphur, when heated together, slowly combine to form carbon disulphide [F].

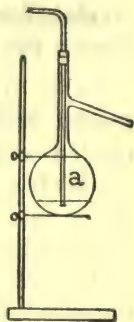


FIG. 74.

5.—To observe the Behaviour of Sulphur when heated.—Place some powdered sulphur in a test-tube and gently heat, observing all that occurs. At about 115° C. it melts, forming an amber-coloured, mobile liquid, which gradually, as the temperature rises, darkens in colour, and becomes less mobile, until, at about 220° to 250° C. it is very dark, and so viscous that the tube may be inverted without loss of sulphur. Further heating causes it to become lighter in colour and less viscous, until its boiling-point, 445° C., is reached. At this temperature it is converted into an orange-coloured gas. On cooling, the same changes occur, but in the reverse order.

Like carbon, this element exists in several “allotropic” forms.

EXPT. 207.—To prepare Octahedral Sulphur.—This is the common form, and has a density of 2.05, and is soluble in carbon disulphide. It is best prepared by

allowing a solution in carbon disulphide slowly to evaporate. Dissolve about 5 gms. of sulphur in carbon disulphide in a large test-tube, loosely close the end with cotton-wool, and allow it to stand. Examine the crystals formed. (How many faces has the crystal?) Sketch the crystal.

EXPT. 208.—To prepare Prismatic Sulphur.—Transparent, needle-shaped crystals are obtained by cooling molten sulphur, with a density of 1.98, which are soluble in carbon disulphide. Melt some sulphur in a pot or crucible, and allow it to cool. As soon as a crust has formed, pierce two holes through it, and pour out the still liquid portion. Remove the crust and examine the crystals. Expose a few crystals and notice that they become opaque, gradually assuming the octahedral form.

EXPT. 209.—To prepare Plastic Sulphur.—This is an elastic, gummy, transparent substance, with a density of 1.95, and is insoluble in carbon disulphide. It is prepared thus: Boil some sulphur in the retort (*a*) (Fig. 75), and allow the molten sulphur to drop into a beaker of cold water (*b*). Observe that this variety slowly becomes opaque on exposure, and loses its elasticity, regaining the octahedral form.

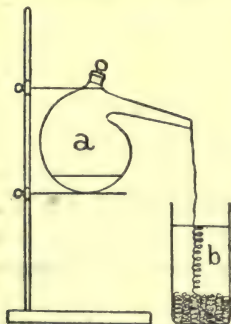


FIG. 75.

EXPT. 210.—To prepare Amorphous Sulphur.—Add dilute hydrochloric acid to a solution of "hypo" [G]. Collect the precipitate, wash and dry it. Examine with a lens; it is non-crystalline. Shake it up with carbon

disulphide; it is **insoluble**. Heat a little; it burns, forming sulphur dioxide.

These **four different physical substances** can all be **mutually converted into each other** by physical processes, yet all, when burned in air or oxygen, yield **sulphur dioxide only**—*i.e.*, they all consist **solely** of the same chemical element.

EXPT. 211.—To show the Identity of the “Allotropic” Forms of Sulphur.—Repeat Expt. 165 on the allotropy of carbon, using samples of each kind of sulphur. You will find that in every case one part by weight of sulphur forms two parts by weight of sulphur dioxide, and that nothing else is produced.

****CHAPTER XIX**

HYDROGEN SULPHIDE

(Sulphuretted Hydrogen)

CHEMICAL formula, SH_2 . Molecular weight, 34. Density, 17 (or 0.00153 gm. per c.cm.).

This offensive-smelling gas occurs in the gases that escape in volcanic regions, in certain "sulphur" springs, and is produced by the decomposition of organic bodies containing sulphur—*e.g.*, egg, etc. It occurs as an impurity in coal-gas.

It is produced, as we learned in Expt. 206, by the direct union of hydrogen and sulphur, and when certain sulphides are reduced in hydrogen, small quantities of this gas are obtained.

EXPT. 212.—To examine the Action of Acids on Sulphides.—Warm as many sulphides as you can with dilute and strong hydrochloric acid and dilute sulphuric acid. Note that the gas evolved (in most cases) has an offensive odour; it is hydrogen sulphide. Thus **metallic sulphides are decomposed by dilute acids liberating hydrogen sulphide.**

EXPT. 213.—To prepare Sulphuretted Hydrogen.—Place in the Woulffe's bottle (Fig. 24, *a*) some ferrous sulphide [H], and pour on to it some hydrochloric (1 : 1) or sulphuric acid (1 : 2), and collect over water the gas evolved. Note that the water has the smell of the

gas. (Why ?) Test the gas as below. (What remains in *a* ?)

EXPT. 214.—To examine the Properties of Hydrogen Sulphide.—Compare this gas with its *analogue*, hydrogen oxide (water). Inspection shows us that the gas is **colourless**, possesses an **offensive odour**, and a **sickly sweet taste**.

N.B.—This gas is **poisonous**, and experiments with it should be conducted in a fume-closet.

1. Place a jar of the gas over water. Observe the steady rise of the liquid. The gas, as we learned above, is **moderately soluble** in cold water. The solution has the smell and reactions of the gas.

2. Test the aqueous solution and the gas with litmus; it is **acid**.

3. Expose a little of the solution to the action of light for some time. Observe that **sulphur [J]** is precipitated and the smell of the gas disappears.

4. Apply a light to a jar of the gas; it **burns [K]** with a blue flame, giving off pungent fumes of sulphur dioxide, but the taper is extinguished. It is a **non-supporter of combustion**.

5. **Caution.**—When you have satisfied your instructor that all air is expelled from the apparatus (How will you test this ?), light the gas as it escapes from the end of the delivery-tube, and make the flame impinge on the surface of a cold evaporating-dish. Observe that sulphur **[J]** is deposited.

6. Bring a jar of sulphur dioxide mouth to mouth with a jar of sulphuretted hydrogen, and observe the precipitation of sulphur **[L]**. (What kind of process is this ?)

7. Bubble the gas through a little acidified potassium bichromate **[M]**, and observe the change in colour. **sulphuretted hydrogen is a reducing agent.**

8. Repeat 7, using potassium permanganate [N] and observe the rapid discharge of the colour.

9. Bubble the gas through strong sulphuric acid [O]. Observe the precipitation of sulphur and the evolution of sulphur dioxide. **Hydrogen sulphide reduces sulphuric acid**, and cannot therefore be dried by this substance.

10. Pass the gas slowly over heated potassium [P], tin, lead, silver, etc., in the apparatus (Fig. 27). Observe that hydrogen escapes, a sulphide of the metal remaining.

11. Pass the gas through the solutions of lead nitrate [Q], mercuric nitrate, copper sulphate, zinc sulphate, cadmium chloride, tin chloride, ferric chloride, etc. Observe that **hydrogen sulphide precipitates many metals as insoluble sulphides**. Notice that with ferric chloride reduction to ferrous chloride first occurs, sulphur being precipitated.

In order to find the density of the gas, **pure, dry hydrogen sulphide** is required. This is obtained by heating antimony sulphide with hydrochloric acid [R], and drying the gas produced by passing it through a U-tube of calcium chloride.

12. Proceed to find the density of the gas as in Expt. 59 (10), and verify the fact that its density is 0.00153 gm. per c.cm. Hence its density, compared with hydrogen, is $\frac{0.00153}{0.00009}$ —i.e., 17.

EXPT. 215. — **To find the Composition of Sulphuretted Hydrogen by Volume.**—Since the gas burns, producing sulphur and water only, and, moreover, as it is formed from hydrogen and sulphur alone, it can consist of these two elements only. We have learned above that the gas is readily decomposed by metals.

Repeat Expt. 130, placing freshly reduced iron filings or granulated tin in the boat [S]. It is found that sulphuretted hydrogen contains its own volume of hydrogen. From this volume composition we may deduce its composition by weight, for—

1 volume of hydrogen sulphide weighing	17
contains 1 volume of hydrogen weighing	<u>1</u>
and hence contains sulphur weighing	16

So that its composition by weight is—

$\frac{1}{17}$, or 5.88 per cent. of hydrogen.

$\frac{16}{17}$, or 94.12 per cent. of sulphur.

EXPT. 216. — To examine Sulphides. — We have already seen that hydrogen sulphide is an acid, and is therefore absorbed by alkalis. (Verify this.) It is such a weak acid, however, that it does not neutralize alkalis (Verify this), but forms hydrosulphides, which are soluble and alkaline to litmus. These hydrosulphides may be used instead of the gas for precipitating the insoluble sulphides. (Compare the alkali hydroxides.) Repeat Expt. 212. Most sulphides are decomposed by acids liberating hydrogen sulphide. Warm various sulphides with nitric acid, potassium permanganate, or bichromate, acidified with hydrochloric acid, test the solutions for sulphates as in Expt. 187, and verify the fact that sulphides are converted into sulphates by oxidizing agents.

SOME COMMON SULPHIDES.

The sulphides are important ores from which metals are obtained.

MERCURIC SULPHIDE occurs in nature as *cinnabar*. It is prepared for use as the pigment *vermilion* by subliming mercury

and sulphur. As obtained by precipitation, it is a black insoluble solid, unaffected by acids.

LEAD SULPHIDE occurs as the mineral *galena*—a dense black lustrous solid. It is obtained as a black precipitate, as in Expt. 214. Nitric acid converts it partially into the sulphate.

COPPER SULPHIDE is a black solid, soluble in nitric acid.

STANNOUS SULPHIDE—a chocolate-coloured solid, and STANNIC SULPHIDE, a bright yellow substance, are obtained by precipitation with hydrogen sulphide. Both are soluble in the caustic alkalies and in ammonium sulphide solution.

FERROUS SULPHIDE, a black solid, is obtained by precipitation or by heating a mixture of iron and sulphur (Expt. 38). It is chiefly used as a source of the gas.

IRON DISULPHIDE occurs as *iron pyrites*. It is a source of sulphur and sulphur dioxide.

ZINC SULPHIDE is found in nature as *blende*. When pure, it is nearly white, and is obtained by precipitation with hydrogen sulphide in an alkaline solution. It is soluble in dilute acids.

CHAPTER XX

NITRIC ACID AND ITS PRODUCTS

CHEMICAL formula, HNO_3 . Molecular weight = 63
[$= 1 + 14 + 3 \times 16$].

This well-known acid does not occur free in nature (Suggest why), although its salts, with sodium, potassium, calcium, and other metals, occur in the soil and plant-juices. Large deposits of sodium nitrate are found in Chili, and potassium nitrate (nitre, saltpetre) exists as an efflorescence on the soil of hot countries. The acid also is found in small quantities in rain after thunderstorms.

EXPT. 217.—To examine the Action of Strong Sulphuric Acid on Nitrates.—Repeat Expt. 186 (11 with small quantities of as many nitrates as you can, warming *gently*. Test any fumes evolved with blue litmus. Drop in a small piece of copper foil. (What happens? What acid is formed?) Do not your experiments prove that all nitrates when warmed with strong sulphuric acid produce nitric acid—a sulphate of the metal is left?

EXPT. 218.—To prepare Nitric Acid.—Heat 40 or 50 gms. of nitre (potassium nitrate), with an equal quantity of strong sulphuric acid in the retort (Fig. 12, a), closed by a *glass* stopper (Why?), its neck being placed in the mouth of a well-cooled flask (b). A pale

yellow liquid distils over. When the distillation is complete, cool the residue somewhat, and pour out the still liquid portion into a dish. On cooling, a colourless, crystalline mass of **potassium hydrogen sulphate** [A] is obtained. Of course, any nitrate could be used—*e.g.*, copper nitrate would give nitric acid and copper sulphate [B], and so on.

On the manufacturing scale, where *cost* is the *consideration*, a greater proportion of nitrate is used, and the cheaper nitrate, sodium nitrate, or Chili saltpetre (Why ?), is employed. The materials are heated in an iron retort lined with clay, the residue being sodium sulphate [C].

EXPT. 219.—**To examine the Properties of Nitric Acid.**—On inspection, the liquid collected in *b* is found to be a pale, yellow, strongly fuming liquid, with a pungent odour.

1. By means of a bellows blow a gentle stream of air through this liquid; a brown gas is expelled, and a colourless liquid is obtained, which **boils at 86° C.**, and has a **density of 1.53 gms. per c.cm.**

2. Into a test-tube of water add *one spot* of this liquid. (*a*) Taste it—it is **sour**; (*b*) add blue litmus—it is **reddened**; (*c*) add a little washing-soda—effervescence occurs; (*d*) add a little magnesium—hydrogen is given off—it is an **acid**.

3. Place in the strong acid a piece of paper, cloth, wood, feather, etc. Note the **yellow stain** on animal and vegetable substances. Wash and test them—their nature is destroyed. **Nitric acid is very corrosive.**

4. Add a few drops of strong nitric acid to some indigo solution—the latter is **bleached**.

5. Heat sawdust on a sand-tray until charring

begins; remove the flame; drop on to it a few spots of the strong acid—it bursts into flame. (What does this show ?)

6. Boil a little sodium sulphite or sulphur with nitric acid and test the solution for a sulphate [D] with barium chloride (Expt. 187).

These experiments show that **nitric acid is an oxidizing agent.**

7. Try the effect of weak and strong nitric acid on as many metals as you can. (Is hydrogen evolved ? Do all metals dissolve ?) Evaporate the liquids down and try to get crystals. (What are they ?)

You will find that nitric acid attacks all metals except gold and platinum—hence its name, **aqua fortis**—generally giving brown fumes and a nitrate. (Tin and antimony form their **oxides**.) (Can you suggest why hydrogen is not liberated when metals react with nitric acid ?)

8. Neutralize nitric acid with as many bases as you can (Expt. 170), and obtain crystalline specimens of the **salts**—viz., **nitrates**.

9. Neutralize nitric acid with as many carbonates as you can (Expt. 170), and obtain the corresponding **nitrates**.

10. Dissolve some of the green crystals—ferrous sulphate—in cold water, and add, slowly, nitric acid. A **brown layer** [E] is formed, which on warming or shaking disappears, a brown gas escaping. This is used as a **test for nitric acid**; hence for **nitrates**.

EXPT. 219A.—To show that Nitric Acid contains **Hydrogen**.—Place some magnesium ribbon in a test-tube, cover it with water, and add to it, spot by spot, nitric acid. A colourless gas is evolved, which explodes on applying a light. It is **hydrogen**, which must

have come from the **acid**, since magnesium is an element; magnesium nitrate [F] remains. Other metals—*e.g.*, iron, zinc, etc.—especially in a finely divided state, also liberate hydrogen from dilute nitric acid.

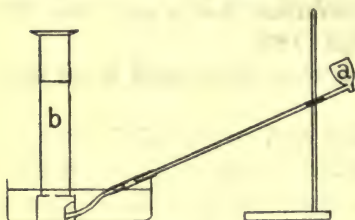


FIG. 76.

EXPT. 220.—To show that Nitric Acid contains Oxygen.—1. Pour strong nitric acid slowly into the bowl of the churchwarden clay pipe, arranged as shown (Fig. 76), having a portion of its stem red-hot. Collect the issuing

gas over water, and after shaking, test with a glowing wood-splinter. This is relighted; the gas is **oxygen** [G].

2. Take any nitrate [H] (except sodium and potassium), which is formed from a metal (an element) and nitric acid, and heat it in a small tube. Test the gas evolved as before; it is **oxygen**. (What is the residue?)

3. Heat potassium or sodium nitrate in a hard-glass tube (Fig. 14), and collect the gas evolved. Examine it and verify the fact that it is oxygen, but observe that no brown fumes are produced in this case. The residue is potassium [I] or sodium *nitrite*. Add a spot of dilute hydrochloric acid, and observe the brown fumes produced.

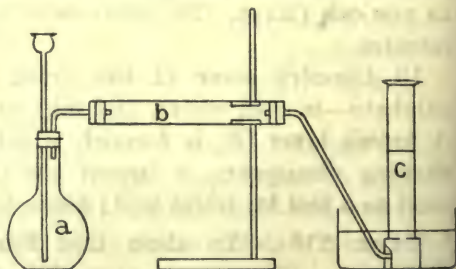


FIG. 77.

EXPT. 221.—To show that Nitric Acid contains Nitrogen.—Into the flask (a) (Fig. 77) place some of the element

copper, and pour on to it some moderately dilute nitric acid. Let the issuing gas pass through the hard-glass tube (*b*) containing red-hot iron filings (freshly reduced), or copper, and collect the escaping gas over water. Test this gas (1) with a lighted taper, (2) with lime-water. The gas is **nitrogen** [**J**]. Examine the residue in *a*. (What is it? What does it show?)

EXPT. 222.—To examine Nitrates—i.e., the Salts—of Nitric Acid.—They are well-defined crystalline solids. Test them with regard to their solubility in water; all are soluble. Repeat Expt. 220 by heating as many nitrates as you can, closely observing any gases evolved; test with a glowing wood-splinter, and examine the residue. All nitrates (except potassium, sodium, and ammonium) when heated yield a brown gas and oxygen, and an oxide of the metal remains. This may be used (1) as a test for nitrates; (2) as a general method for preparing metallic oxides—e.g., lead nitrate [**K**] gives oxygen, nitrogen peroxide, and leaves lead oxide; copper nitrate [**H**] gives oxygen, nitrogen peroxide; and leaves copper oxide; and so on. Sodium and potassium nitrate [**I**] slowly liberate oxygen, whilst the *nitrites* of these metals remain. For the action of heat on ammonium nitrate, see Expt. 231.

Expt. 216 taught us that all nitrates reacted with strong sulphuric acid to form nitric acid and sulphates, and the former is identified by its action on copper, or its formation of a brown layer with cold ferrous sulphate solution. These are therefore used as tests for nitrates. Try them with various nitrates.

Since nitrates so readily liberate oxygen, they are vigorous oxidizing agents.

EXPT. 223.—To show that Potassium Nitrate is an Oxidizing Agent.—1. Melt a little nitre in a porcelain

dish, and drop small fragments of hot carbon into it. The carbon burns vigorously. (What is formed ?) When quite cold, add dilute hydrochloric acid to a little of the residue, and test the gas with lime-water. (What is the residue ?)

2. Repeat 1, but drop in sulphur. (What is produced in this case ?)

This oxidizing action explains why hydrogen is not set free when metals and nitric acid interact. The hydrogen first set free—a reducing agent—is oxidized, by the further nitric acid present, ere it can escape.

The Source of Nitrates in the Soil (Nitrification) (see p. 135).—Apart from the mineral deposits of nitrates (Chili, etc.), these salts are found in all soils, appearing as an incrustation on the soil of hot countries, and form the only source of nitrogen for most plants. Small quantities are produced in the atmosphere during thunderstorms, etc., but far larger supplies are furnished, through the agency of certain micro-organisms present in the soil, from nitrogenous organic matter. Animal matter on decay produces ammonia, which is converted by one of these bacteria into nitrites; the nitrites are then transformed into nitrates by a second germ. Thus, the cycle of nitrogen between plants and animals is completed. The nitrates are extracted from the soil by treatment with water, and crystallization. This process is imitated in Europe by exposing heaps of animal refuse, slaked lime, and garden soil to the air at intervals, moistening the heaps with stable drainage. The resulting calcium nitrate is then dissolved out as above.

SOME COMMON NITRATES.

POTASSIUM NITRATE (NITRE, SALTPETRE) is found in certain soils. It is made from the commoner and less valuable sodium nitrate by adding potassium chloride. When the solution is concentrated, common salt is precipitated and the solution crystallized. It may also be obtained as in Expt. 170.

Nitre forms large, well-shaped crystals, and is very soluble. Unlike the sodium salt, it is not deliquescent, and, being a

powerful oxidizer (Expt. 223), it is used in making gunpowder, fireworks, etc. When heated, it loses oxygen, and forms potassium nitrite, and it is decomposed by sulphuric acid, nitric acid being formed.

SODIUM NITRATE (CUBIC NITRE, CHILI SALTPETRE) occurs in large quantities in Chili. It is used in the manufacture of nitre and nitric acid, and was prepared in Expt. 170.

Sodium nitrate behaves exactly like the potassium salt, save that it crystallizes in cubes, is much more soluble, and deliquescent.

AMMONIUM NITRATE, a white crystallized solid, is prepared in Expt. 230, and its behaviour when heated is examined in Expt. 231. When placed in water, a considerable lowering of temperature is brought about, and it is often made use of as a mild "freezing-mixture."

MERCURY NITRATES.—When a large quantity of mercury is shaken with a small amount of dilute nitric acid, a solution is obtained which, on evaporation by exposure, yields colourless soluble crystals of *mercurous nitrate*. *Mercuric nitrate* is obtained by heating a little mercury with strong nitric acid. It is a colourless, crystalline, soluble solid.

LEAD NITRATE is a soluble crystalline solid, obtained by dissolving the metal, its oxide, or carbonate in nitric acid (Expt. 170). The action of heat on this substance is examined in Expt. 227.

Other nitrates are prepared and behave like lead nitrate with respect to heat and strong sulphuric acid. *Copper nitrate* (blue) is very corrosive and deliquescent. *Silver nitrate* (*lunar caustic*) stains the skin black, and is often used for marking linen. It is the reagent for testing for chlorides (*q.v.*).

Oxides of Nitrogen.

Nitric Oxide.—Chemical formula, NO. Molecular weight, 30 ($= 14 + 16$). Density = 15 (or 0.00135 gm. per c.cm.).

In Expt. 221 a colourless gas was formed, which we found to contain nitrogen and oxygen. We will prepare the gas and examine it more closely.

EXPT. 224.—To prepare Nitric Oxide.—Place some copper wire or foil in the Woulffe's bottle (Fig. 24, *a*), provided with a thistle funnel and delivery-tube, and pour on to it nitric acid diluted with its own bulk of water. Collect the gas liberated over water. (What does this show?) The gas collected may have a distinct brownish colour, which disappears on standing or being shaken with water, when a colourless insoluble gas is left. When the action is over, filter the liquid and evaporate it to a small bulk, allow it to cool and crystallize. Blue crystals of **copper nitrate** [**J**] result.

A *purer* product is obtained by warming together a mixture of nitric acid (or a nitrate), sulphuric acid [**L**], and ferrous sulphate, collecting the gas over water.

EXPT. 225.—To study the Properties of Nitric Oxide.—As we learned above, this gas is **colourless**, and **insoluble in water**.

1. Expose a jar of the gas to the air; a *brown* gas is produced—**nitrogen peroxide** [**M**]. This property **distinguishes the gas from all other gases**.

2. Place some *feebly burning* phosphorus (**caution**), sulphur, taper, sodium, etc., into the gas; they are **extinguished**. Now plunge the same **brightly burning** phosphorus into the gas; it continues to **burn**, forming **phosphorus pentoxide** [**N**] and **nitrogen**.

3. Bubble the gas through a cold solution of ferrous sulphate; the gas is absorbed, forming a **brown solution**. (Compare the brown layer test for nitrates.)

Thus ferrous sulphate may be used to *absorb*, and so to *estimate* the amount of, nitric oxide in a gaseous mixture.

Pure dry nitric oxide is obtained from the gas prepared from nitric acid and ferrous sulphate above, which is *dried* by passing it through a **U-tube** contain-

ing calcium chloride (Why not sulphuric acid? see p. 219), and collected over mercury.

4. Introduce into a jar of the dry gas over mercury a little neutral litmus solution; it remained unaffected—*i.e.*, nitric oxide is **neutral**.

5. **To find the Density of the Gas.**—Repeat Expt. 59 (10), filling the flask with pure dry nitric oxide, and verify the statement that its **density is 0.00135 gm. per c.cm.**, or its **relative density** compared with hydrogen is $\frac{0.00135}{0.00009}$ —*i.e.*, 15.

EXPT. 226.—To show that Nitric Oxide contains Nitrogen and Oxygen.—Repeat Expt. 221. Nitrogen collects, while iron or copper oxide [J] remains in the hard-glass tube, thus showing the presence of these two elements.

This may also be used to find its composition by weight. (Suggest an apparatus, etc.)

Nitrogen Peroxide.—This gas is, as we learned above, formed when nitric oxide is exposed to air or oxygen, and when nitrates are heated.

EXPT. 227.—To prepare Nitrogen Peroxide.—1. Take a large test-tube and divide it into three equal parts by india-rubber rings or pieces of gummed paper, and fill two-thirds, as above, with pure nitric oxide. Take a smaller tube and mark on it a volume equal to one-third of the

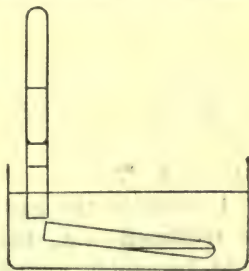


FIG. 78.

larger tube. Fill this with pure oxygen, and transfer it quickly to the larger one, over water, as shown (Fig. 78), and observe (1) a brown gas is formed; (2) its volume is

equal to that of the nitric oxide; (3) that it *all rapidly dissolves*. Thus, two volumes of nitric oxide and one volume of oxygen form two volumes of nitrogen peroxide [M].

2. Heat *dry* powdered lead nitrate [K] (other nitrates would do) in the tube (a) (Fig. 79), and pass the gases evolved through a U-tube immersed in a freezing-mixture of ice and salt. Lead

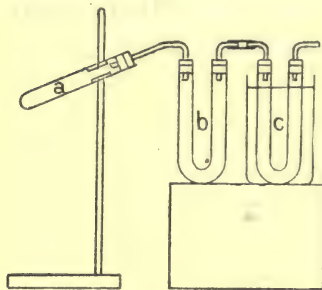


FIG. 79.

oxide remains, oxygen escapes; whilst a yellowish-green liquid collects in the U-tube. It is **liquid nitrogen peroxide**. If the gas is required, fit a delivery-tube to the U-tube, close the other limb with a cork, and remove the U-tube from the freezing-mixture. As the U-tube

becomes warmer, the liquid evaporates, forming the gas. This liquid has a molecular weight of 92, its chemical formula being N_2O_4 .

EXPT. 228.—To study the Properties of Nitrogen Peroxide.—This is a **brown** gas, which, as found above, is **soluble in water**, and possesses a characteristic odour and acid taste. It **readily liquefies**, as is shown by Expt. 227.

1. Bubble the gas through blue litmus solution; it is **reddened**—i.e., the gas is **acidic**. If the water is cold, two acids, nitrous [P] and nitric acids, are formed; but if hot water is employed, nitric acid [Q] only is produced, whilst nitric oxide escapes.

2. Shake a jar of the gas with a little mercury; the mercury is *oxidized* [R], whilst the colour of the gas

is destroyed, nitric oxide being formed. **Nitrogen peroxide is a vigorous oxidizing agent.**

3. Plunge some **brightly burning** substances—phosphorus (**Danger**), sulphur, sodium, etc.—in the gas. They continue to **burn**, forming **oxides [S]**, whilst **nitrogen remains—i.e., nitrogen peroxide supports combustion, but is incombustible.**

4. Repeat Expt. 59 (10), drying the gas by passing it through a **U-tube** of calcium chloride (Fig. 32). The **density** is found to be relative to hydrogen, not 46, as expected, but *about 38*.

When the **density** is determined at **140° C.**, it is curious to note that it is **23**. What does this indicate? It would seem that, as this gas is heated, its relative density decreases, and with it a marked change in the colour takes place, which shows that the substance represented by the formula N_2O_4 is broken down into the substance having the formula NO_2 . The reverse change occurs on cooling (see Expt. 235).

EXPT. 229.—To show that Nitrogen Peroxide contains Nitrogen and Oxygen.—This is borne out by the first method of preparation, and if Expt. 221 is repeated, **nitrogen** collects, whilst copper or iron **oxide [T]** is produced.

Nitrous Oxide.—Chemical formula, N_2O . Molecular weight, $44 = [2(14) + 16]$. Density = 22 (*i.e.*, 0.00198 gm. per c.cm.).

This was discovered by Priestley in 1776, and was found by Davy to have the curious effect of causing “laughter,” if inhaled mixed with air or oxygen; hence the common name of “*laughing*” gas. If inhaled pure, it produces brief insensibility, and is used for this purpose by dentists.

EXPT. 230.—To prepare Ammonium Nitrate.—Half fill a dish with dilute nitric acid and add ammonium

hydrate, drop by drop, stirring and testing with litmus, after each addition, until neutral. Evaporate the neutral solution to a small bulk; cool, and set aside to crystallize. A white crystalline solid ammonium nitrate [**U**] is formed.

EXPT. 231.—To examine Ammonium Nitrate.—1. Put some of the crystals in a small quantity of water, and observe the fall in temperature; this forms a mild freezing-mixture.

2. Warm a little of the solid with caustic soda solution; an alkaline gas [**V**], with a characteristic odour (ammonia) is evolved.

3. Heat a few crystals gently in a small tube; observe the steam evolved. Pass the gases through water, and collect in a test-tube. A colourless gas is obtained, which relights a glowing ember and soon dissolves. Now heat more strongly; an explosion results, and a brown gas is produced [**W**]. (What is this ?)

EXPT. 232.—To prepare Nitrous Oxide.—Gently heat about 30 gms. of ammonium nitrate in the flask (Fig. 51, *a*), fitted with a delivery-tube, and collect the gas evolved over warm water. (Why ?) If the gas is required pure, it is first passed through a solution of ferrous sulphate (What does this remove ?), and then dried by passing through a U-tube containing calcium chloride, the dry gas being collected over mercury.

EXPT. 233.—To study the Properties of “ Laughing ” Gas.—Observation teaches us that this gas is a colourless and practically odourless gas.

1. Place a jar of the gas, mouth downwards, in cold water; the water rises—*i.e.*, the gas is soluble in water.

2. Test with litmus; no effect is observed—it is neutral.

3. Plunge a glowing wood ember into a jar of the gas; it is rekindled—the gas is a **vigorous supporter of combustion**, and an **incombustible**.

4. **Danger**.—Plunge some burning phosphorus [X], etc., in the gas; burning continues, **oxide of phosphorus** and **nitrogen** resulting.

5. Its **density** may be found, as in Expt. 59 (10), to be 0.00198 gm. per c.cm., or its **relative density** is 22.

To distinguish Nitrous Oxide from Oxygen.—There is a danger of confusing this gas with oxygen. The following experiments may be used to distinguish between these two gases:

1. Nitrous oxide dissolves in cold water, oxygen does not.

2. Place a jar of nitrous oxide mouth to mouth with a jar of nitric oxide, and remove the plates; no brown fumes are produced, such as are formed when oxygen and nitric oxide are mixed.

It should be noticed that nitrous oxide produces insensibility, and then death, whereas oxygen is a supporter of life.

EXPT. 234.—To show that Nitrous Oxide contains Nitrogen and Oxygen.—Repeat Expt. 221; nitrogen escapes, and iron oxide [Y] is formed.

(Suggest *one* experiment, in detail, to find the composition of the air or of any of the oxides of nitrogen.)

***EXPT. 235.—To determine the Composition by Volume of the Oxides of Nitrogen**.—Since we have discovered that these compounds can be reduced by substances fond of oxygen, their composition can be easily established. Repeat Expt. 130, placing some freshly reduced iron filings in the porcelain boat. Iron oxide remains in the boat in each case, and we find—

1. Nitrous oxide contains its own volume of nitrogen.
2. Nitric oxide contains half its volume of nitrogen.
3. Nitrogen peroxide contains from a *half to three-quarters* of its volume of nitrogen. (Why this curious result in this case ?)

Since the densities of nitrogen, oxygen, and these oxides are known, we may deduce the composition by weight of the first two.

1 volume of nitrous oxide weighing	22
contains 1 volume of nitrogen weighing	14
and hence contains oxygen weighing	<u>8</u>

So that nitrous oxide contains

$$\frac{14}{22} \text{ i.e., } \frac{7}{11}, \text{ or } 63\cdot63 \text{ per cent. of nitrogen by weight;}$$

$$\frac{8}{22} \text{ i.e., } \frac{4}{11}, \text{ or } 36\cdot37 \text{ per cent. of oxygen by weight.}$$

Similarly, nitric oxide contains

$$\frac{7}{15}, \text{ or } 46\cdot67 \text{ per cent. of nitrogen by weight;}$$

$$\frac{8}{15}, \text{ or } 53\cdot33 \text{ per cent. of oxygen by weight.}$$

Nitrogen peroxide contains

$$\frac{7}{23}, \text{ or } 30\cdot44 \text{ per cent. of nitrogen by weight;}$$

$$\frac{16}{23}, \text{ or } 69\cdot56 \text{ per cent. of oxygen.}$$

These results may be readily confirmed by finding their compositions by weight *directly*, as in Expts. 111 and 221. The nitrogen may be collected over water or in a vacuum globe.

****CHAPTER XXI**

AMMONIA

(Spirits of Hartshorn, Volatile Alkali)

CHEMICAL formula, NH_3 . Molecular weight, 17 ($= 3 + 14$). Density, 8.5 (or 0.000765 gm. per c.cm.).

This pungent-smelling gas has been met with incidentally in the preceding sections (Expt. 231), and may be produced in small quantities by sparking nitrogen and hydrogen. It is present in small amount in the air, and is produced by the decay of animal nitrogenous matter. It may be obtained by distilling such substances as hair, horn, coal, etc. Certain metals—*e.g.*, magnesium—when heated in air or nitrogen, absorb the latter, forming *nitrides*, which are decomposed by water liberating ammonia gas and forming the metallic oxide.

EXPT. 236.—To examine the Action of Alkalies on Ammonium Salts.—Warm small quantities of as many ammonium salts as you can with caustic potash solution, caustic soda solution, lime, lime and water, soda-lime, etc. Test the gas evolved (1) by smell, (2) with litmus-paper, and observe that a **pungent-smelling, alkaline gas** is obtained in every case. Thus, **ammonium salts are decomposed by any alkali, ammonia gas being liberated.** (What remains in the above experiments ?)

EXPT. 237.—To prepare and collect Ammonia Gas.—Place in the flask (a) (Fig. 80) some ammonium chloride (or any other salt of ammonium), an equal bulk of lime [A] (or any other alkali) and just sufficient water to make

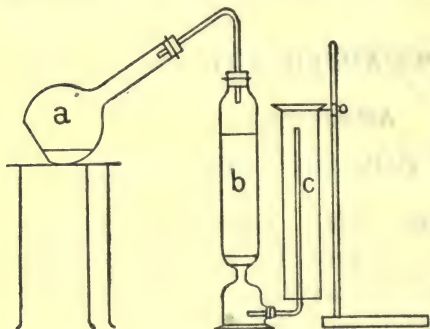


FIG. 80.

the whole into a paste. Since the gas is alkaline, it is dried by passing it through the tower (b), containing *solid* caustic soda, soda-lime, etc., and collected by displacing air downwards. (What property does this show? Why not

dry it with calcium chloride? See below.) The gas may also be collected over mercury. (What advantage has this method? What is the residue in a? Verify your answer.)

In Expt. 161 we discovered that when coal was destructively distilled, an aqueous ammoniacal liquor was obtained. This liquid, obtained in enormous quantities at the gas-works, is the source of all the ammonia in commerce. It is treated with excess of lime and distilled, and the various salts are obtained as below. Owing to the inconvenience of storing the gas, a strong solution in water—the “liquor ammonia” of the laboratory—is prepared.

EXPT. 238.—To prepare Ammonia Solution.—The gas, prepared as above, is passed through a series of wash-bottles containing water, as shown (Fig. 81). *a* contains a *small quantity of water*. (Why?) The best

and strongest solution is obtained in *b*, then *c*, and so on. Such a solution contains 700 volumes of the gas in one volume of water, or 35 per cent. by weight, and has a sp. gr. of 0.88. It is a convenient source of the gas in the laboratory, since by gently warming large volumes of the gas are obtained.

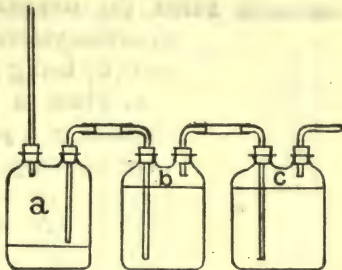


FIG. 81.

EXPT. 239.—To examine the Properties of Ammonia.—As we have

already discovered, ammonia is a **colourless, pungent-smelling, alkaline gas**. It is less dense than air (Expt. 237), and **very soluble in water** (Expt. 238). The latter fact is well illustrated by the “*fountain*” experiment.

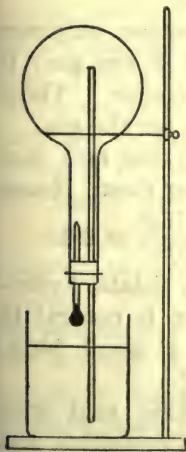


FIG. 82.

1. Fill the flask (Fig. 82) by displacement, with the gas, and fit it with a cork provided with a long tube and a “*filler*” containing a little water. Place the open end of the long tube in a dish of water tinted with red litmus, and squeeze the bulb of the “*filler*.” The gas rapidly dissolves in the water, rising in a fountain, and the litmus is turned blue.

2. Plunge a taper into a jar of ammonia, and observe that the gas does **not support combustion**, nor does it burn itself. Ammonia can, however, be made to burn thus:

3. Warm some strong ammonia solution in *a* (Fig. 83)

At the same time pass oxygen into the wide tube (b) containing cotton-wool at c, and apply a light. The ammonia burns (in oxygen) with a characteristic greenish-yellow flame, water and nitrogen [B] being produced.

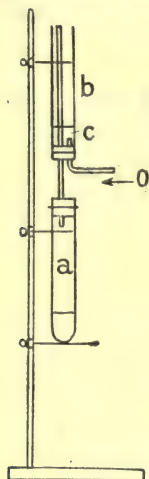


FIG. 83.

4. Place a few lumps of calcium chloride in a jar of the dry gas standing over mercury. Observe the rise of the mercury in the jar. Hence calcium chloride [C] absorbs (and unites with) ammonia, and cannot therefore be used to dry the gas.

5. Place a glass rod dipped in strong hydrochloric acid into a jar of the dry gas; dense white fumes of ammonium chloride [D] are produced.

6. Hold a drop of copper sulphate solution on a glass rod in the gas; it acquires a fine blue colour. These (5 and 6) are useful tests for the gas.

7. Repeat Expt. 59 (10), and verify the fact that this gas has a density of 0.000765 gm. per c.cm. Hence its density relative to hydrogen is $\frac{0.000765}{0.00009}$, or 8.5.

8. Bubble ammonia gas into some dilute hydrochloric acid solution until the product is neutral to litmus. Evaporate the solution to a small bulk; crystallize, and examine it as below [D].

9. Neutralize some dilute hydrochloric acid with ammonia solution, as in Expt. 230. Crystallize, and compare the product [D] with that obtained in 8; thus: (1) heat—both sublime; (2) warm with caustic soda solution—both yield ammonia [E]. (Are they the same substance?) Expts. 8 and 9 may be repeated

with other acids, and it is found that **ammonia gas** and **ammonia solution both neutralize acids to form salts.**

Ammonium.—Compare a solution of ammonia with a solution of caustic soda or potash; thus: (1) with litmus, (2) with copper sulphate solution [F], (3) with acids [G], and note that ammonia solution acts exactly like sodium and potassium hydroxides. For these reasons we "*believe*" ammonia solution is a hydroxide formed by the union of ammonia with water, and it is called **ammonium hydroxide [H]**, or **ammonium hydrate**, although all attempts to isolate "ammonium" have hitherto failed.

EXPT. 240.—To show the Presence of Nitrogen and Hydrogen in Ammonia.—Since ammonia is formed when a mixture of hydrogen and nitrogen is subjected to the action of electric sparks, these elements alone are present in the gas. This may also be easily shown thus: Pass a stream of dry ammonia gas through the hard-glass tube (*a*) containing dry copper oxide, and send the product through a tube (*b*), kept cold, as shown in Fig. 67. Any residual gas can be collected over water and examined. Note the copper oxide is reduced [I] to copper. Water collects in *b* (hence the presence of hydrogen in ammonia), whilst nitrogen collects. This experiment may be arranged to **find the composition of ammonia by weight.** (How would you arrange it for this purpose, and how would you carry out the experiment in order to insure an accurate result?) In this way **ammonia** is found to **contain—**

$\frac{14}{17}$, or 82.36 per cent. of nitrogen by weight,
 $\frac{3}{17}$, or 17.64 per cent. of hydrogen by weight.

EXPT. 241.—To find the Composition of Ammonia by Volume.—When ammonia gas is sparked for some time, its volume is found to increase until, on cooling, it is exactly twice the volume of the original gas. On applying a light to this gas, an explosion results, for the ammonia has been decomposed into nitrogen and hydrogen. This may be utilized to find its composition.

1. Into the eudiometer (Fig. 41) pass a few c.cms. of dry ammonia gas, and spark it until no further increase in volume is observed. (What is now in the eudiometer?) Cool it, adjust the pressure and read the volume. Now admit a large volume of pure oxygen, and proceed as in Expt. 89. A contraction is found, and since one volume of oxygen and two volumes of hydrogen unite to form liquid water [B], two-thirds of this contraction is hydrogen and one-third is oxygen. Calculate thus:

Volume of ammonia	= 3.4 c.cms.
After sparking, volume of nitrogen and hydrogen	= 6.8 „
Volume of oxygen, nitrogen, and hydrogen	= 29.7 „
Volume after explosion	= 22.0 „

Hence 3.4 c.cms. of ammonia formed $\frac{2}{3}(29.7 - 22.0)$; i.e.,
5.1 c.cms. of hydrogen;
and 3.4 c.cms. of ammonia formed $6.8 - 5.1$; i.e.,
1.7 c.cms. of nitrogen.

So that 3.4 c.cms. of ammonia contain 5.1 c.cms. of hydrogen and 1.7 c.cms. of nitrogen; i.e.,

2 volumes of ammonia contain 3 volumes of hydrogen and 1 volume of nitrogen.

2. This may be verified thus: It will be seen later (Expt. 250) that one volume of chlorine combines with one volume of hydrogen to form hydrogen chloride, and we make use of this fact in the following experiment. Fill

the burette, fitted with a tap-funnel as shown (Fig. 84), with chlorine (see Expt. 255), and divide it into three equal volumes by rubber bands. Place in the tap-funnel some strong ammonia solution, and allow it to enter spot by spot. At first a greenish flame is produced; when the colour of the chlorine is no longer visible, run in a small excess of ammonia, close t_1 and open t_2 under water. Observe that the water rushes in and fills two-thirds of the burette. Now three volumes of chlorine have disappeared, and hence have united with *three volumes of hydrogen*. Test the residual gas; it is *nitrogen*. So that **three volumes of hydrogen unite with one volume of nitrogen to form ammonia**.

From these results, knowing the relative densities of these gases—viz., 1, 14, and 8.5—we may *deduce* its composition by weight, for 3 volumes of hydrogen weighing 3 unite with 1 volume of nitrogen weighing 14, to produce 2 volumes (see 1) of ammonia weighing $2 \times 8.5 = 17$. Hence ammonia consists of

$\frac{3}{17}$ hydrogen and $\frac{14}{17}$ nitrogen by weight,

thus confirming the result obtained in Expt. 239.

EXPT. 242.—To examine Ammonium Salts.—As we have already learned, these may be obtained by neutralizing acids with ammonia, or “ammonium” hydroxide. Examine as many of these salts as you can, and verify the following statements:

1. All ammonium compounds are volatile; some sublime, others decompose.
2. All ammonium salts are soluble.
3. All, when warmed with alkalis, liberate ammonia gas.



FIG. 84.

****The Solubility of Gases in Water.**

Our study of gases in the preceding sections enables us to draw some conclusions on the solubility of gases in water.

1. In the first place, it is obvious that the **solubility depends upon the nature of the gas**, for 1 c.cm. of water at 0°C . and under standard pressure dissolves 0.02 c.cm. of hydrogen, 0.05 c.cm. of oxygen, 0.02 c.cm. of nitrogen, 1.80 c.cms. of carbon dioxide, 79.8 c.cms. of sulphur dioxide, 503 c.cms. of hydrogen chloride, 1,148 c.cms. of ammonia. A study of these and other experimental facts reveals the following interesting points:

(1) Those gases which **dissolve most readily** are also **most readily liquefied**.

(2) Whilst the **neutral gases are only slightly soluble**, the gases that are **acidic or alkaline** are much **more soluble**, and the **more acidic or alkaline** a gas is, the **more soluble** it is.

Now, in many cases we have evidence that *acidic and alkaline gases enter into chemical union with water*. For instance, when hydrogen chloride, ammonia, etc., dissolve in water, considerable heat is evolved—a good sign of chemical union. Again, carbon dioxide and sulphur dioxide form acids by union with water; they are acid anhydrides (Expt. 74, 3).

2. The second point which we discover is that gases dissolve more readily in cold water than in hot water. In Expt. 232 we made use of this in order to collect nitrous oxide, for at standard pressure 1 c.cm. of water at 0°C . dissolves 1.30 c.cms. of this gas; at 10°C ., 0.92 c.cm.; at 20°C ., 0.67 c.cm. Moreover, in Expt. 54

we found that **all air** could be **expelled** from water by **boiling**, and this is **true** in *almost all cases*. Hydrogen chloride is an exception to this rule, for when a strong solution of this gas is boiled, it loses hydrogen chloride until a particular strength is reached. We may therefore say that—

(1) **The solubility of all gases decreases as the temperature rises.**

(2) **The whole of the gas in solution may be expelled in almost all cases.**

Since the volume of a gas is so greatly affected by pressure, it is of importance to determine the effect of pressure on the solubility of gases. Henry, who investigated this, found, in 1803, that **unit volume of a liquid dissolves the same volume of a gas at all pressures, or the mass of a gas dissolved by unit volume of a liquid is proportional to the pressure.** This is known as **Henry's law.**

CHAPTER XXII

HYDROCHLORIC ACID GAS—HYDROGEN CHLORIDE AND ITS PRODUCTS

(Spirits of Salt—Muriatic Acid)

CHEMICAL formula, HCl . Molecular weight, 36.5 ($= 1 + 35.5$). Density, 18.25 (or 0.00164 gm. per c.cm.).

This acid does not occur free, but is very plentiful as its salts, especially as sodium chloride—common salt or rock-salt—both as a mineral and in sea-water. Many other chlorides, notably of potassium, magnesium, etc., are found in enormous quantities.

EXPT. 243.—To study the Effect of Strong Sulphuric Acid on Chlorides.—Repeat Expt. 186 (10), warming as many chlorides as you can with strong sulphuric acid; all give off steamy acid fumes of hydrochloric acid gas, whilst a sulphate of the metal remains. This may be used therefore—

1. As a test for chlorides. (Mention another.)
2. As a general method for preparing hydrogen chloride, since all chlorides with strong sulphuric acid liberate hydrochloric acid and form sulphates.

EXPT. 244.—To prepare Hydrogen Chloride.—Common salt, or sodium chloride, is generally used, although any other chloride would do, and, to avoid frothing, this is *fused*, or *rock-salt* is employed. This is covered with strong sulphuric acid [A] in the flask (Fig. 68, *a*), and

gently warmed; the gas is collected, after washing with a little water in *b*, by displacing air upwards. (What does this show? What can this gas be dried with?)

If the gas is required **dry**, strong sulphuric acid is used in *b*, or the gas is passed through a U-tube containing calcium chloride, and the dry gas is collected over mercury. When the action is completed and the liquid in *a* is cooled, it solidifies to a crystalline mass of **sodium hydrogen sulphate**.

On the large scale a further quantity of salt is used, and a much higher temperature, when **sodium sulphate [B]** remains.

Owing to the inconvenience of storing the gas, a strong solution is prepared in water—the hydrochloric acid of the laboratory.

EXPT. 245.—To prepare Hydrochloric Acid (Solution).—The gas, prepared as in the previous experiment, is bubbled through a series of wash-bottles containing water, as in Expt. 238. Such a solution contains about 450 volumes of gas in each volume of water, or 41 per cent. by weight, and has a density of 1.20. This solution is therefore a very convenient source of the gas in the laboratory, the water being removed by a dehydrating agent—*e.g.*, strong sulphuric acid.

EXPT. 246.—To obtain the Gas from the Solution.—Strong sulphuric acid is placed in the tap-funnel (Fig. 85) and is allowed to drip slowly into strong hydrochloric acid. A copious stream of hydrogen chloride escapes.

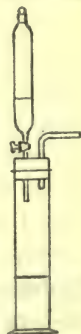


FIG. 85.

EXPT. 247.—To study the Properties of Hydrogen Chloride.—As we have already learned, this substance is a colourless, fuming, pungent, acid gas.

1. Hold a piece of blue litmus-paper in the gas; it is reddened—*i.e.*, it is **acid**.

2. It is **denser than air** (roughly, one and a half times), as is shown by its mode of collection, and also by the following experiments:

(a) Pour the gas from one jar into a jar of air, and test with litmus.

(b) Find its **density**, as in Expt. 59 (10); it is **0.00164 gm. per c.cm.**, so that its **relative density**, compared with hydrogen, is $\frac{0.00164}{0.00009}$ —*i.e.*, **18.25**.

3. Plunge a lighted taper into the gas; the flame is extinguished, and the gas does not burn—*i.e.*, it is **incombustible**, and a **non-supporter of combustion**.

4. As we have seen above, the gas is **very soluble in water**. This is admirably shown by the "*fountain*" experiment. Repeat Expt. 239 (1).

5. Bring a jar of dry ammonia gas mouth to mouth with a jar of dry hydrogen chloride; clouds of the white solid, **ammonium chloride [C]** are produced.

As Expt. 57 taught us, **dilute hydrochloric acid dissolves metals, liberating hydrogen, and forming its salts—chlorides—**whilst in Expt. 170 we found that it **neutralizes bases, decomposes carbonates, etc., forming chlorides**.

We now proceed to examine hydrochloric acid gas, with the object of finding its composition.

EXPT. 248.—To show the Presence of Hydrogen in Hydrogen Chloride.—1. Place some zinc [**D**] or magnesium in dilute hydrochloric acid, and test the gas evolved with a lighted taper; an explosion results. The gas is **hydrogen**.

2. Pass dry hydrogen chloride over heated Zn, aluminium, iron filings, or sodium [**E**], and collect the

gas liberated over water (Fig. 77). A colourless gas collects, which explodes when a light is applied—viz., **hydrogen**. (What remains in the hard-glass tube ?)

EXPT. 249.—To show the **Presence of Chlorine**.—

1. Warm some of the strong acid with manganese dioxide, and examine the gas liberated; it is a disagreeable-smelling gas, with a greenish colour—hence its name, **chlorine** [F].

2. Pass dry hydrochloric acid gas over heated manganese dioxide [F] in a hard-glass tube; observe the formation of **chlorine** and **water**.

These experiments show that hydrogen and chlorine are components of hydrogen chloride, but they do not satisfy us that only these elements are present.

EXPT. 250.—To show that no other Element is present in Hydrogen Chloride, and to determine its Composition—1. **Synthesis**.—Take the apparatus shown

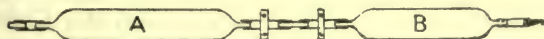


FIG. 86.

(Fig. 86), and find the volume of each part. *B* is less than *A*. When dry, fill the tube *B* with dry chlorine and *A* with dry hydrogen, keeping the clips closed. When full, close each with a plug, invert, and open the clips. Expose the mixture to daylight until the colour of the chlorine disappears. Open one end of the apparatus under mercury, and observe that *nothing enters or leaves* the apparatus. Now open the end under water tinted with litmus; this enters and fills the greater part of the apparatus. Find the volume of the gas left, and test the gas. It will be found that **one volume of chlorine and one volume of hydrogen unite to form two volumes of hydrogen chloride** [G].

2. Analysis (Electrolysis).—Fill with strong hydrochloric acid solution, the U-tube (A) (Fig. 87), provided with two carbon electrodes, and arrange to collect any

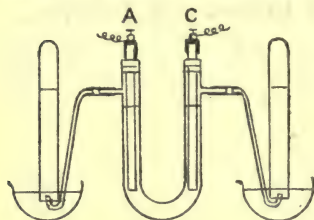


FIG. 87.

gas liberated over strong brine. Pass an electric current through the liquid, and observe that **equal volumes** of gas collect in each tube. The gas over the **anode**, on testing, proves to be **chlorine**; the other, over the **cathode**, **hydrogen** [H],

which confirms the results obtained by synthesis.

Now that we know only hydrogen and chlorine are present, we may find the composition by volume from its action on metals.

EXPT. 251.—To determine the Composition of Hydrogen Chloride.—1. Repeat Expt. 130, using freshly reduced iron filings, zinc [E], or aluminium, etc. Chloride of the metal results, whilst hydrogen is liberated, and it is found that **hydrogen chloride contains half its volume of hydrogen**.

Caution.—2. Fill over mercury a tube, divided into two equal parts by means of india-rubber bands, with dry hydrogen chloride. Insert, by means of crucible-tongs, a small piece of sodium [E]. A slight explosion occurs, and the mercury rises and fills half the tube; the residual gas, on being tested, proves to be hydrogen, thus verifying the conclusion above. The tube of hydrogen chloride may be filled by displacement and placed upright in a dish of sodium amalgam, with a like result, although the action is more slow.

From the composition by volume so found the **composition by weight** can be *deduced*, for 2 volumes of

hydrogen chloride weighing 2 (18·25), or 36·5, contain 1 volume of hydrogen weighing 1, and therefore chlorine weighing 35·5. So that **hydrogen chloride consists of**

$$\frac{1}{36\cdot5}, \text{ or } 2\cdot75 \text{ per cent. of hydrogen by weight,}$$

$$\frac{35\cdot5}{36\cdot5}, \text{ or } 97\cdot25 \text{ per cent. of chlorine by weight.}$$

As we have seen in the above experiment, hydrogen readily unites with chlorine, even without applying a light—*i.e.*, even more readily than it unites with oxygen.

EXPT. 252.—To show that Hydrogen burns in Chlorine.—1. Lower a lighted jet of hydrogen into a jar of chlorine; it continues to burn, forming dense fumes of hydrochloric acid gas.

2. Half fill a jar with chlorine and half with hydrogen by displacement over brine; wrap a duster around it, and apply a light. A violent explosion occurs, comparable with that of hydrogen and oxygen, showing that *hydrogen* [G] and *chlorine* have very great affinity for each other.

In a similar way, just as **hydrogen** will reduce metallic oxides, forming water and the metals, so it will reduce chlorides of the metals, forming the metals and hydrogen chloride.

EXPT. 253.—To reduce Silver Chloride in Hydrogen.—Repeat Expt. 59 (11), using silver chloride in the boat, and passing the escaping gas through blue litmus. When all air is expelled, heat it, observe that silver is produced, and that the litmus is turned red, owing to the hydrochloric acid gas [I] formed.

This could be used to find (1) the **composition of metallic chlorides**; (2) the **composition of hydrogen**

chloride by weight. (Suggest suitable apparatus, and mention any precautions you would take in performing these experiments.)

Chlorine.

Chemical symbol, Cl. Atomic weight, 35.5. Molecular weight, 71. Density, 35.5 (or 0.00317 gm. per c.cm.).

This gas was first discovered by Scheele, 1775, by treating *muriatic acid* with the mineral *pyrolusite*—manganese dioxide. In 1810 Sir H. Davy proved it to be an element, and gave it the name **chlorine**, from its *greenish* colour.

Caution.—Chlorine gas is very injurious, and should not be prepared in the open laboratory, but in a fume cupboard.

EXPT. 254.—To study the Behaviour of Oxidizing Agents on Hydrochloric Acid.—Warm strong hydrochloric acid with small quantities of various oxidizing agents—*e.g.*, nitric acid, potassium chlorate, manganese dioxide, lead peroxide, etc.—and test with litmus-paper the gas evolved; note the colour and odour. Do your experiments justify the statements—

1. Hydrochloric acid with oxidizing agents yields chlorine?

2. Metallic peroxides and hydrochloric acid liberate chlorine and form chlorides?

Owing to the discovery of this gas by the oxidation of hydrochloric acid and to the prevalence of Lavoisier's oxygen theory of acids, chlorine was thought to be an oxide; hence the name "*oxymuriatic acid*." Davy, in 1808, began a course of research into the nature of the gas, which occupied him until 1810. By no means whatever could he abstract oxygen from this gas; neither electric sparks nor strongly ignited carbon affected

it. Phosphorus and sulphur failed to give their oxides, whilst the gas passed over oxides of potassium, barium, and other metals, produced salts and expelled all the oxygen in the oxides. He thus conclusively demonstrated the elementary nature of chlorine, and overthrew Lavoisier's erroneous theory of acids.

A mixture of three parts of hydrochloric acid and one part of nitric acid is known as **aqua regia**, since it will dissolve the royal metals, gold and platinum. Its solvent action is due to the liberation of **nascent chlorine** [J], which rapidly attacks the metals, by forming soluble chlorides.

Since, as we have learned in Expt. 243, all chlorides, when warmed with strong sulphuric acid, yield hydrochloric acid, we can prepare chlorine from all chlorides, **for all chlorides and strong sulphuric acid with a metallic peroxide will thus liberate chlorine, and form sulphates.**

EXPT. 255.—To prepare Chlorine from Hydrochloric Acid.—Place in the flask (Fig. 68, *a*) some metallic peroxide—usually manganese dioxide—and cover it with concentrated hydrochloric acid. On being warmed, a greenish gas is evolved, which, after drying by being passed through strong sulphuric acid in *b*, is collected by displacing air upwards. (What does this show?) When the action is completed, filter the liquid in *a*, evaporate it to a small bulk and cool. Pink crystals of **manganese chloride** [K] remain.

EXPT. 256.—To prepare Chlorine from Chlorides.—Use the same apparatus as in the previous experiment. Place some chloride—*e.g.*, sodium chloride and manganese dioxide—in the flask. Pour down the thistle funnel some strong sulphuric acid. Chlorine is collected, whilst sulphates of the metals remain—*e.g.*, in this case **sodium sulphate and manganese sulphate** [L].

EXPT. 257.—To study the Properties of Chlorine.—This element is observed to be a **greenish-coloured** gas, with a **pungent, suffocating odour**. It is **injurious if inhaled**, and is **denser than air**, as is shown by the method of collection above.

1. Pour chlorine from one jar into a lower jar. Its density, found as below, shows that it is **2·5 times** that of the air.

Chlorine is a very active element, combining readily with metals and non-metals.

2. Plunge a lighted taper into a jar of the gas; it continues to burn with a lurid flame, fumes of **hydrochloric acid gas [G]** being formed, and clouds of **soot (carbon)** deposited.

3. Lower a lighted jet of **hydrogen** into the gas; the hydrogen burns, forming **hydrogen chloride [G]**.

4. Place a filter-paper, soaked in **turpentine** and gently warmed, in a jar of the gas. It bursts into flames, producing **hydrogen chloride [M]** and clouds of **soot**.

Thus, **chlorine has a great affinity for hydrogen**. As we have seen in Expt. 250, the mixed gases unite slowly in daylight, exploding violently if a light is applied, whilst chlorides are reduced by hydrogen (Expt. 253). This is further illustrated by its action on hydrogen sulphide.

5. Pass some chlorine into a jar of **hydrogen sulphide [N]**. Observe the **sulphur** deposited, whilst fumes of **hydrogen chloride** fill the jar.

6. Introduce a sheet of **Dutch metal** (thin brass) into the gas; it burns spontaneously.

7. Drop some powdered **antimony [O]** into chlorine; spontaneous combustion occurs, **antimony trichloride** being formed. Add water; a white liquid is obtained. Test the liquid with litmus; it is acid.

8. **Danger.**—Place a fragment of **phosphorus** on a deflagrating spoon and plunge it into a jar of chlorine. The phosphorus melts and burns spontaneously, forming **phosphorus trichloride** or **pentachloride** [P]. Add water and litmus; an acid solution is obtained.

N.B.—Non-metallic chlorides are decomposed by water forming acids.

Sulphur and carbon are unaffected; arsenic behaves like antimony.

9. Melt a piece of sodium in a hard-glass tube (Fig. 27), and pass the gas through it. The **sodium** burns brilliantly, producing **sodium chloride** [E].

10. Repeat this experiment, using **quicksilver**; a flame plays over the surface of the metal, and a sublimate of **mercuric chloride** [Q] is produced.

Other metals behave in a similar manner.

Thus, chlorides may be prepared from the elements as well as from hydrochloric acid.

Chlorine is often used as a **disinfectant** ("chloride of lime," or "bleaching-powder" being employed), for it destroys the organisms producing putrefaction, etc. It is also a strong bleaching agent.

EXPT. 258.—To study the **Bleaching Action of Chlorine**.—1. Bubble chlorine through water; a liquid which possesses the colour and odour of the gas is obtained, termed **chlorine-water**. Thus, **chlorine is soluble in water**.

2. Half fill a bottle with a saturated solution of chlorine-water, and fill up with chlorine; close it securely, and expose it to sunlight. When the colour has disappeared, open the bottle and examine the contents. It no longer smells of chlorine, and the gas, on testing, proves to be **oxygen**, whilst the liquid is **hydrochloric acid solution**—i.e., **chlorine decomposes**

water, forming hydrogen chloride, and liberating oxygen [R], so that wet chlorine is an oxidizing agent.

Verify this (1) by shaking some lead oxide in water and bubbling chlorine through the mixture. Examine the puce-coloured solid (i.) by heating, (ii.) by warming with concentrated hydrochloric acid. (Is it lead peroxide [S]?) (2) Bubble the gas through sulphur dioxide [T] solution, and observe that the odour of sulphur dioxide disappears. Test the solution for a sulphate. Sulphuric and hydrochloric acids are formed.

3. Bubble the gas through litmus or indigo solution; the colour is discharged—*i.e.*, **chlorine bleaches**.

4. Write with writing-ink on a printed label, and lower it into a jar of chlorine; the ink is bleached, but not the printing-ink, which contains charcoal.

5. Carefully *dry* a jar of the gas with lumps of calcium chloride, etc., and place in it some dry turkey-red twill. No bleaching occurs. Now moisten the cloth; bleaching takes place at once. Thus we see that **bleaching by chlorine only occurs in the presence of moisture**. As we have seen above (2), chlorine and water react, liberating oxygen. This oxygen is in a **very active state** (the *nascent*—*i.e.*, new-born, or **atomic**—condition), and it is this active oxygen that **bleaches by oxidation [R]**. Compare the bleaching by sulphur dioxide, which is by reduction or simple combination.

****EXPT. 259.—To prepare "Bleaching Liquor."**—Fill a test-tube with chlorine by displacement, and invert it over some strong solution of caustic soda or potash. Observe that the gas is dissolved or absorbed. Now prepare a larger amount of this solution by bubbling the gas through a solution of caustic potash until it escapes freely, and examine it as below:

1. Add litmus or indigo solution. (Does it bleach ?)
2. Shake a little lead oxide with this solution. (Does it oxidize ?)
3. Add some dilute acid. (Does chlorine escape ?)
4. Boil the remainder of the liquid, and test again, as before. It seems to have lost its bleaching and oxidizing powers, nor does it liberate chlorine when warmed with acids. Evaporate it until crystals appear. (What are they ?) Test their behaviour when heated. (What is this substance ?) Evaporate the "mother liquor," and examine the solid obtained by (1) warming with strong sulphuric acid (What is evolved ?), (2) dissolving a little in water and adding some solution of silver nitrate. (What is this solid ?)

Thus, **chlorine is absorbed by cold alkali solution, forming a "bleaching liquor" [U], which, by boiling, is converted into a chlorate and chloride.**

****EXPT. 260.—To make "Bleaching-Powder."**—Pass chlorine up through a funnel containing some fresh, carefully slaked lime, and examine it as below :

1. Mix it with water; smell, and test it with wet litmus-paper. (Does it contain chlorine, and does it bleach ?)
2. Add some dilute acid, and warm. (What do you learn ?)
3. Warm some lead oxide with this powder and water. (Is it oxidized to lead peroxide ?)

Thus, a **bleaching-powder** has been formed—it is, in fact, **chloride of lime [V]**, or common **bleaching-powder**.

EXPT. 261.—To find the Density of Chlorine.—The gas, prepared as in Expt. 255, and dried by passing through strong sulphuric acid or calcium chloride, is passed into the apparatus (Fig. 26) for half an hour or so, and then weighed. **Its density is calculated as in**

Expt. 59 (10), and found to be 0.003195 gm. per c.cm. So that its density, compared to that of hydrogen, is $\frac{0.003195}{0.00009}$, or **35.5**, whilst compared to air, it is $\frac{0.003195}{0.001293}$ —i.e., **2.5**.

Should a jar of **pure, dry** gas be required, it may be collected over paraffin oil, but not mercury (Why?); or by displacing air upwards.

The compounds formed by the union of the elements with chlorine are termed **chlorides**, and are identical with the chlorides formed from hydrochloric acid.

EXPT. 262.—**To examine Chlorides.**—Test as many chlorides as you can, and determine (1) whether they are soluble in water or not; (2) their effect on litmus; (3) the effect produced when silver nitrate is added to their solution, acidified with nitric acid; (4) the effect produced by warming them with strong sulphuric acid; and (5) with manganese dioxide and strong sulphuric acid. Verify the following conclusions:

Chlorides are well-defined crystalline salts, mostly soluble in water. Their solutions yield a white precipitate of silver chloride with silver nitrate [W] solution (Why?); whilst they all liberate hydrochloric acid gas when warmed with strong sulphuric acid and chlorine, if manganese dioxide is also present.

SOME COMMON CHLORIDES.

SODIUM CHLORIDE (COMMON SALT).—This well-known substance, all-important to man, occurs very widely as rock-salt and in sea-water, etc. It has been prepared from its elements by the action of the metal on hydrogen chloride, and from its hydroxide, and carbonate.

EXPT.—**To prepare "Pure" Salt from Rock-Salt, etc.**—Prepare a hot saturated solution of the impure salt and divide it into three portions:

1. Cool rapidly, and observe that few, if any, crystals appear. Examine the solubility curve of common salt (p. 56), and note that the solubility of sodium chloride is almost constant at all temperatures.

2. Add an equal bulk of strong hydrochloric acid.

3. Pass hydrogen chloride (gas) into the solution until no more precipitate is formed. Compare your results in 2 and 3. Why do you obtain so much more **pure precipitated salt** in 3 than in 2?

Sodium chloride crystallizes in small cubes, is soluble, and has its characteristic salty taste. It is the main source of sodium and its compounds, and also of hydrogen chloride, and indirectly of chlorine. It is also used to prepare insoluble chlorides by double decomposition.

POTASSIUM CHLORIDE is very similar to sodium chloride, and is prepared similarly.

CALCIUM CHLORIDE is chiefly obtained as a by-product in the preparation of carbon dioxide and ammonia (Expts. 126 and 237). It may also be prepared by the action of hydrochloric acid on lime or slaked lime.

Dry chloride of calcium is a valuable drying agent in the laboratory, and is deliquescent. Its solution yields colourless crystals.

MERCURIC CHLORIDE has been prepared from its elements, and is a dense crystalline, not very soluble, solid. It may also be obtained by dissolving mercuric oxide in aqua regia. It is intensely poisonous, and a valuable antiseptic.

MERCUROUS CHLORIDE, an insoluble, dense, white solid, is obtained by precipitation from mercurous nitrate. When heated, it breaks down into mercury and mercuric chloride.

MAGNESIUM CHLORIDE is a deliquescent, crystalline solid. When heated, its crystals leave a residue of magnesium oxide.

STANNOUS CHLORIDE is obtained by dissolving tin in hydrochloric acid, and forming soluble crystals. It is a powerful reducing agent.

STANNIC CHLORIDE is obtained by passing chlorine over heated tin. It is a colourless, fuming liquid, which, with water, forms colourless, soluble crystals.

LEAD CHLORIDE, a white, insoluble solid, soluble in hot water, is prepared by double decomposition (Expt. 177). It may also be obtained by boiling lead with hydrochloric acid.

IRON CHLORIDES.—1. **FEROUS CHLORIDE** is produced as a green, very soluble salt when iron dissolves in hydrochloric acid. It may also be prepared by heating iron in dry hydrogen chloride gas. Ferrous chloride readily undergoes oxidation, and is therefore a reducing agent.

2. **FERRIC CHLORIDE** may be made by heating iron in excess of chlorine, or by oxidizing ferrous chloride with nitric acid. It is a very soluble, deliquescent, crystalline solid.

COPPER CHLORIDES.—1. **CUPRIC CHLORIDE**, a greenish-blue crystalline solid, is obtained when copper burns in chlorine, or when copper oxide is dissolved in hydrochloric acid. It is reduced by stannous chloride, or by boiling with excess of copper to—

2. **CUPROUS CHLORIDE.**—This may also be obtained by dissolving the cuprous oxide in strong hydrochloric acid, and diluting the solution. It is a white, insoluble solid.

SILVER CHLORIDE, an insoluble white solid, obtained by double decomposition. It melts to a horny mass (it occurs as “horn silver”), and is easily reduced (Expt. 253).

Combustion.

When we speak of “combustion,” or burning, we usually refer to the more common occurrences—viz., where the combustion takes place in air or oxygen. Our experiments, however, show us that combustion can take place in other gases besides these, for in Expt. 128 we observed that magnesium and potassium burned in carbon dioxide, a gas which is usually considered to be a typical **non-supporter of combustion**. Iron and copper burn in sulphur vapour (see Expts. 38 and 206), while many brightly burning substances continue to burn in the oxides of nitrogen (see Expts. 225, 228, and 233), and the active gas chlorine causes spontaneous combustion of metals, non-metals, and hydrogen compounds (see Expts. 257).

We thus see that combustion, or burning, has a much wider meaning than that commonly given to it.

In all the cases discussed above vigorous chemical union occurs, accompanied by light and heat—*i.e.*, often by a **flame**. If we slow down the chemical action, the flame ceases, and often all light disappears; we do not then refer to the union as a combustion. We may therefore define **combustion (or burning)** as **rapid chemical union, accompanied by heat and light (often by a flame)**, while the production of light may often, and a flame must always, be regarded as a visible sign of chemical union.

It should be noticed that we may have light and heat without chemical union—*e.g.*, an electric glow-lamp, red-hot platinum, etc. It will also be observed that a **flame** is only produced when **gases** are concerned in the action. When light and heat are produced, but no flame, it is termed **incandescence**—*e.g.*, limelight, incandescent gas-mantles, etc.—and this may or may not denote chemical combination.

EXPT. 263.—**To show that Oxygen burns in Hydrogen.**—As we learned in Expt. 86,

a jet of hydrogen burns more brightly in a jar of oxygen than in air, and we said that the former substance was combustible, whilst the latter was a supporter of combustion. Pass a slow stream of oxygen from a gas-holder through a right-angled tube. Light a jar of hydrogen at its mouth, and bring it over the tube as shown (Fig. 88). Observe that the **oxygen burns [X]**—*i.e.*, is a **combustible**—and the **hydrogen is here a supporter of combustion**, while water is deposited on the sides of the gas-jar.

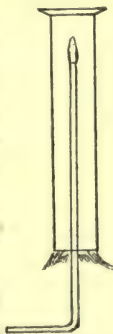


FIG. 88.

EXPT. 264.—**To show that Chlorine burns in Hydrogen.**—Repeat the former experiment, supplying

chlorine [G] instead of oxygen. Here again the **supporter of combustion**, chlorine, now becomes the **combustible**, and *vice-versa*.

EXPT. 265.—To show that Air burns in Coal-Gas.—Join the right-angled tube (a) to the coal-gas supply, and fill the apparatus with coal-gas.



Light it at *b*; here we have “**coal-gas**” **burning in the air**. Now push up *c* a lighted taper; observe the smoke drawn up into the lamp-glass—i.e., air is drawn into the apparatus. After one or two attempts, a flame will appear at *d*. Here we have just the reverse of what occurs at *b*, for here “**air**” **burns in coal-gas**.

From these experiments it is evident that the terms “combustible” and “supporter of combustion” are merely **relative**, for both are burning, and both are supporting combustion, each of the other.

It is for convenience only that we say the substance **surrounding** a flame is the **supporter of combustion**, while the **substance inside** the flame is called the **combustible**.

CHAPTER XXIII

CHEMICAL LAWS

What is a Chemical Law? The question at once arises, Do the elements unite to form compounds in a haphazard way, or in a definite and regular manner? In Expt. 38, on the formation of compounds, we found that the iron used up all the sulphur it required, while the excess merely burned away. (Why?) Can the iron only combine with a certain amount of sulphur? Again, in Expt. 89 we found that, although a large excess of hydrogen was exploded with oxygen, the latter **always** united with **exactly** twice its volume of hydrogen, and any more than this remained unaffected at the end of the experiment. When determining the composition of water by weight (Expt. 88), the same fact stood out prominently, for many **different experimenters**, using quite **different substances** and **methods**, all found the **same results**—viz., one part by weight of hydrogen **always** united with **exactly** eight parts by weight of oxygen. Even when, as in the case of carbon and sulphur, different “physical forms” were used, this truth was confirmed. So that, without doubt, **elements combine together in a perfectly regular manner by weight (and volume).**

This behaviour of elements, which we have discovered by experiment—viz., what they can or cannot do—seems to point out that elements are **governed** in

some way, and when we express the general result of many experiments, we express the **laws** which they obey in their chemical reactions. In Chapter X. we made the acquaintance of the first important natural law in Chemistry, which was first definitely stated by Lavoisier as a result of his experiments—viz., **the law of conservation of matter—i.e., “the total quantity of matter taking part in any chemical change remains constant,”** or, more briefly, **“matter cannot be created or destroyed.”** We cannot hope to **prove** this law absolutely, for new changes appear every day; but the nature of the evidence is conclusive—viz., that, while all known experiments agree with, and therefore **illustrate** the law, **no** experiment is known which **disobeys** or **contradicts** it.

The second chemical law has also been met with in many of the preceding experiments, for we have made the same substance—*e.g.*, copper oxide, water, carbon dioxide, sulphur dioxide, and many other compounds—in many **different ways**, beginning with **different materials**, and have **always found their composition the same**, whether we examined them by synthesis or analysis. This is stated as **the law of constant composition, or the law of definite proportions—“the same compound is always composed of the same elements, united in the same proportion by weight”**—*i.e.*, having once found that a particular compound is made up of certain elements united in a certain relation by weight, we shall **always** know its composition, however it is obtained. Refer here to our experiments on water, the allotropy of carbon and sulphur, and see how this law has been proved. In order to **verify** this law further, any compound may be selected; it may be made in any way; it may be analyzed in any way, etc., as below:

EXPT. 266.—To verify the Law of Constant Composition for Copper Oxide.—1. Heat a weighed amount of clean copper filings in a weighed boat in a hard-glass tube (Fig. 27) in a stream of oxygen [A], until, on cooling and weighing, no further change in weight occurs, and calculate the percentage of copper and oxygen in copper oxide, thus :

Constant weight of boat and oxide = 3.72 gms.

Weight of boat and copper = 3.48 „

∴ Gain = weight of oxygen = 0.24 gm.

Weight of boat and copper = 3.48 gms.

Weight of boat alone = 2.54 „

∴ Weight of copper = 0.94 gm.

Weight of copper oxide = 0.94 + 0.24 = 1.18 gms.

Percentage of oxygen = $100 \times \frac{0.24}{1.18} = 20.1$ per cent.

Percentage of copper = $100 \times \frac{0.94}{1.18} = 79.9$ per cent.

2. Dissolve a weighed amount of pure copper foil in nitric acid in a weighed dish with a cover-glass, and evaporate the liquid *carefully* to dryness. When near the end, wash the cover-glass into the dish; copper nitrate now remains. Dry gently, and finally heat strongly until only black copper oxide [B] remains; cool and weigh, and calculate as above.

3. Dissolve in a beaker a weighed amount of copper foil in nitric acid and add excess of sodium hydrate solution [C], boiling until the blue hydrate of copper is converted into black copper oxide. Filter off the black oxide, as in Expt. 43, and wash well with hot water until the filtrate is no longer alkaline to litmus-paper. Dry and weigh the oxide, calculating your results as before.

4. Weigh a porcelain boat; place in it about 1 gm.

of pure dry copper oxide [D], and reweigh. Reduce it in hydrogen or coal-gas, as described in Expt. 59, and work out your results as above.

Now compare your four results thus:

Percentage of—			Percentage of—		
1. Copper	=	79.9	Oxygen	=	20.1
2. Copper	=	80.3	Oxygen	=	19.7
3. Copper	=	79.7	Oxygen	=	20.3
4. Copper	=	80.5	Oxygen	=	19.5
Average	=	80.1	Average	=	19.9

Thus, allowing for the errors in your experiments, these results show that copper oxide is **always** made up of 80.1 per cent. of copper and 19.9 per cent. of oxygen (the correct values are 79.75 per cent. of copper, and 20.25 per cent. of oxygen)—i.e., is of constant composition.

EXPT. 267.—To verify the Law of Constant Composition for Silver Chloride.—1. Weigh a clean porcelain boat; place in it about 1 gm. of thin silver foil [E], and reweigh. Fit up the apparatus as shown in Fig. 27, and pass dry chlorine through the tube, in a draught cupboard, heating strongly meanwhile. When all metallic appearance has disappeared, cool and reweigh. Reheat in a stream of chlorine, cool and reweigh until of constant weight. Work out your results thus:

Constant wt. of boat and silver chloride = 5.29 gms.

Weight of boat and silver = 4.96 „

∴ Gain = weight of chlorine = 0.33 gm.

Weight of boat and silver = 4.96 gms.

Weight of boat alone = 3.96 „

∴ Weight of silver = 1.00 gm.

Percentage of silver = $100 \times \frac{1.00}{1.33} = 75.19$ per cent.

Percentage of chlorine = $100 \times \frac{0.33}{1.33} = 24.81$ per cent.

2. Weigh out accurately about 1 gm. of silver foil; place it in a beaker with a cover-glass and stirring-rod, and add a small quantity of dilute nitric acid [**F**]. Warm until dissolved, and add a solution of hydrochloric acid or common salt until no more precipitate is formed. Boil and filter through two filter-papers of equal weight, as in Expt. 43. Wash the residue on the filter-paper with small quantities of hot water until the filtrate no longer gives a precipitate with silver nitrate solution. Dry and weigh the silver chloride, and calculate as above.

3. Weigh a clean porcelain boat, add about a gramme or two of pure dry silver chloride [**G**], and reweigh. Reduce this in a stream of hydrogen, as in Expt. 59; cool and weigh the silver obtained, and compare your results thus:

<i>Percentage of—</i>			<i>Percentage of—</i>		
1.	Silver	= 75.19	Chlorine	=	24.81
2.	Silver	= 75.05	Chlorine	=	24.95
3.	Silver	= 75.24	Chlorine	=	24.76
Average = 75.16			Average = 24.84		

(Correct results = 75.26 per cent. of silver, and 24.74 per cent. of chlorine.)

Thus, again, the compound silver chloride always has the same composition. Similar experiments with any other compounds always lead to the same result, so that we are justified in believing this law to be true, and the chemist takes it as a sound basis for his future work.

CHAPTER XXIV

EQUIVALENTS

(Combining Weights or Combining Proportions)

SINCE, as we have learned in the last chapter, elements combine together in a definite manner by weight, it is of great importance for the chemist to know the weights in which the elements unite. We might from our experiments draw up a table showing the amounts of each element that unites with one part by weight of oxygen (*I.*), or with eight parts by weight of oxygen (*II.*), or with one part by weight of hydrogen (*III.*), or any other chosen weight of any particular element:

	<i>I.</i>	<i>II.</i>	<i>III.</i>
Carbon ..	0.375	3.00	3.00
Sulphur ..	1.000	8.00	16.00
Magnesium ..	1.500	12.00	—
Sodium ..	2.875	23.00	—
Iron ..	1.875	14.00	—
Copper ..	3.875	31.50	—
Phosphorus ..	0.775	6.20	10.33
Oxygen ..	—	—	8.00
Chlorine ..	—	—	35.50

Such weights are termed “combining weights,” or “proportions”—*i.e.*, the weights or proportions in which the elements combine together.

It is obvious that in order that different chemists may be able to understand one another's results, each

must state what unit is used, or all must use the same unit; and the latter method is the more convenient. Since **hydrogen** is the least dense of all elements, the unit weight of hydrogen (1 gm., 1 ounce, 1 lb., etc.) is **chosen** as the unit of measurement in the determination of combining weights, and we can therefore define it as—

The combining weight of an element is the amount of that element which will combine with unit weight of hydrogen—*i.e.*, the combining weight of hydrogen is 1.

EXPT. 268.—To determine the Combining Weight of Oxygen.—Repeat Expt. 88, calculating thus:

Weight of tube and oxide before = 27.65 gms.

Weight of tube and oxide after = 26.08 „

∴ Loss = oxygen = 1.57 gms.

Weight of receiver and water = 31.23 gms.

Weight of receiver alone = 29.46 „

∴ Gain = water = 1.77 gms.

Weight of hydrogen = water - oxygen = 1.77 - 1.57 = 0.20 gm.

Thus, 0.20 gm. of hydrogen unites with 1.57 gms. of oxygen;

1 gm. of hydrogen unites with $\frac{1.57}{0.20} = 7.85$ gms. of oxygen.

∴ Combining weight of oxygen is 7.85.

(More accurate experiments show that one part of hydrogen combines with eight parts by weight of oxygen, or the combining weight of oxygen is 8).

EXPT. 269.—To find the Combining Weight of Chlorine.—Weigh a porcelain boat containing about 2 gms. of silver chloride [A], and place in the hard-glass tube (Fig. 40, *d*), attached to a weighed U-tube (*e*), containing soda-lime or caustic soda. Pass pure

dry hydrogen through the apparatus until all air is expelled; heat the chloride, and when the reduction is complete, cool and reweigh. (What is the residue ?)

RESULTS—

Weight of boat, etc., before = 5.64 gms.

Weight of boat, etc., after = 4.89 „

∴ Loss = chlorine = 0.75 gm.

Weight of *e* after = 28.55 gms.

Weight of *e* before = 27.78 „

∴ Gain = hydrogen chloride = 0.77 gm.

Weight of hydrogen = hydrogen chloride - chlorine =
 $0.77 - 0.75 = 0.02$ gm.;

So that 0.02 gm. of hydrogen unites with 0.75 gm. of chlorine;

1 gm. of hydrogen unites with $\frac{0.75}{0.02} = 37.5$ gms. of chlorine.

Combining weight of chlorine = 37.5 (accurately 35.5).

Unfortunately, but few elements combine readily with hydrogen, so that our first definition is very much restricted, and we must amplify it so that it applies to a larger number. In the experiments above we have found that 8 parts by weight of oxygen and 35.5 parts by weight of chlorine *each combine with the same quantity* (unit weight) *of hydrogen—i.e., they play the same part* so far as their power of combining with hydrogen is concerned, and these *two quantities of oxygen and chlorine are thus equivalent* to one another, and we speak of these weights—*i.e., combining weights—as equivalents.*

Now, it is obvious that anything which combines with 8 of oxygen or 35.5 of chlorine will be equivalent to 1 of hydrogen, which also does this. So that **equivalents combine with equivalents.**

Again, in the above experiments the metals are left as residues in the boats, and if these quantities are weighed, it will be found that 8 of oxygen left 31.5 of copper, in order to unite with 1 of hydrogen—*i.e.*, the *equivalent of hydrogen* has replaced the *equivalent of copper*; or, as in the second case, the equivalent of silver for 35.5 of chlorine—*i.e.*, the equivalent of chlorine—left 108 of the metal in order to combine with 1 of hydrogen. We may therefore say that **equivalents are replaced by the equivalent of hydrogen.**

In Expts. 79 and 80 we found that 12 gms. of magnesium united with 8 gms. of oxygen, forming 20 gms. of magnesium oxide; whilst Expt. 66 showed us that steam passed over heated magnesium loses its hydrogen, and forms the same compound. In this experiment, if 12 gms. of magnesium are used, we obtain 20 gms. of magnesium oxide and 1 gm. of hydrogen. Here the *equivalent of magnesium* replaces the *equivalent of hydrogen*. We may now make our definition perfectly general; thus—

Definition.—The equivalent of an element is that amount which combines with or replaces unit weight of hydrogen or the equivalent of some other element.

To find the Equivalents of the Metals by Replacement of Hydrogen.

A. From Dilute Acids.

EXPT. 270.—To find the Equivalent of Magnesium. (Zinc, Iron, etc.).—Use the apparatus described in Expt. 62. Half fill the flask with dilute hydrochloric acid, and attach a delivery-tube leading into a trough of water. Weigh out about 0.2 gm. of magnesium [B], and place it in the small tube (*a*), and fit up the appar-

atus, taking care that the acid does not reach the metal. Test the apparatus and make it air-tight, and invert a gas-jar of water over the exit-tube. Lower the metal into the acid, and collect the hydrogen given off. When the action is complete, place a piece of gummed paper at the water-level, remove the jar, fill up to the mark with water, and measure it in a graduated cylinder. From the volume of hydrogen calculate its weight [Expt. 58 (10) showed us that 1 litre, or 1,000 c.cms., of hydrogen weighs 0.09 gm.], and from your results determine the equivalent of magnesium, thus:

RESULTS—

Weight of magnesium = 0.20 gm.

Volume of hydrogen = 185 c.cms.

Weight of hydrogen = $\frac{185}{1000} \times 0.09 = 0.01665$ gm.

So that 0.01665 gm. of hydrogen is replaced by 0.20 gm. of magnesium;

1 gm. of hydrogen is replaced by

$$\frac{0.20}{0.01665} = 12.01 \text{ gms. of magnesium.}$$

Hence the **equivalent of magnesium = 12.01.**

Using this apparatus, slight errors occur. (What are they? How would you avoid them?)

Repeat the experiment, using dilute sulphuric acid [C], and show that the *equivalent is independent of the acid used.*

EXPT. 271.—To find the equivalent of a Metal—e.g., Zinc.—Use the same flask, etc., as above, but attach it to an aspirator, as shown (Fig. 63), and make the whole air-tight; when this is so, no water will flow from *e*. Half fill the small flask with weak acid; weigh

about 0.5 gm. of zinc [D], and put it in the small tube, arranging as before. Attach it to the aspirator again, and when air-tight, place a graduated cylinder under the exit-tube. Lower the metal into the acid; the action commences, and water collects in the cylinder. When the action is over (How will you know?), allow the apparatus to cool (Why?), measure the volume of water expelled—i.e., the volume of hydrogen set free—and calculate as before.

Are all the errors in Expt. 270 avoided in this? Suggest an experiment in which the weight of the hydrogen could be determined directly.

B. From Alkalies.

We have learned that certain metals liberate hydrogen when warmed with alkalies (see Expt. 85), and hence this reaction may be used to determine equivalents. The apparatus in Expt. 270 could not be employed (Why?), but that of Expt. 271 yields good results.

EXPT. 272.—To find the Equivalent of Aluminium (Zinc, etc.).—Weigh out 0.20 gm. of aluminium and place it in a hydrogen-tube, between two pieces of copper gauze, as shown (Fig. 90). Attach a delivery-tube and arrange to collect the hydrogen [E] liberated over water. Two-thirds fill a boiling-tube with sodium hydrate (or potassium hydrate) solution, and heat nearly to the boiling-point. Lower the metal into the hot soda, and collect the gas set free. Measure its volume as in Expt. 270, and calculate its equivalent. (What errors are there in this experiment?)

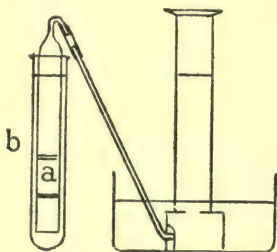


FIG. 90.

EXPT. 273.—To determine the Equivalent of Aluminium.—Repeat Expt. 271, placing cold caustic soda solution in the flask, and adding about 0.3 gm. of aluminium. When air-tight, warm the flask and collect the water expelled. After all the aluminium has dissolved, allow the whole apparatus to become quite cold, with the end of the delivery-tube under the water; observe the volume and calculate as before.

C. From Water.

Metals, like sodium and potassium, which decompose water, may have their equivalent determined in a similar way.

EXPT. 274.—To find the Equivalent of Sodium.—**Caution.**—Sodium, etc., must not come into contact with anything wet or damp. Weigh a small dry ignition-tube, place in it a piece of sodium, and re-weigh. Place this tube and contents between the copper gauze in a dry hydrogen-tube, and immerse in a boiling-tube of water. When the apparatus is ready, gently tapping the hydrogen-tube will cause the water to enter the ignition-tube, and the hydrogen [F] will be liberated, which may be collected and measured as before.

To find the Equivalents of the Elements by their Union with Oxygen—i.e., to find the weight of the element that unites with eight parts by weight (i.e., the equivalent) of oxygen.

A. By Direct Union.

EXPT. 275.—To find the Equivalent of Copper.—Weigh a clean crucible, place in it about 1 gm. of pure copper powder [G] free from oil or grease, and re-

weigh. Heat strongly, and when the oxidation is complete (How will you know?), cool and reweigh. Calculate as below:

RESULTS—

Constant wt. of crucible and oxide = 17.51 gms.

Weight of crucible and copper = 17.15 „

∴ Gain = weight of oxygen = 0.36 gms.

Weight of crucible and copper = 17.15 gms.

Weight of crucible alone = 15.73 „

∴ Weight of copper = 1.42 gms.

So that 0.36 gm. of oxygen unites with 1.42 gms. of copper;

1 gm. of oxygen unites with $\frac{1.42}{0.36}$ gms. of copper,

and 8 gms. of oxygen unite with

$$\frac{1.42}{0.36} \times 8 = 31.55 \text{ gms. of copper.}$$

Since 8 is the equivalent of oxygen, therefore the equivalent of copper = 31.55.

EXPT. 276.—To find the Equivalent of Lead (Copper, Tin, etc.).—Weigh a short hard-glass tube (Fig. 35) containing a plug of dry asbestos and a clean porcelain boat (*a*). Place in the latter about a gram or two of lead [**H**], and reweigh the whole. Pass oxygen through the apparatus, and heat the tube and contents strongly for half an hour or so. Cool and weigh; reheat in the oxygen stream, cool and reweigh until of *constant weight*. Calculate as above, and compare your results with those obtained in Expt. 275.

EXPT. 277.—To find the Equivalent of Carbon.—Repeat Expt. 129, which was to find the composition of carbon dioxide [**I**] by weight. Here we found that

0.3 gm. of carbon united with 0.82 gm. of oxygen, so that since

0.82 gm. of oxygen unites with 0.30 gm. of carbon;

1 gm. of oxygen unites with $\frac{0.30}{0.82}$ gm. of carbon;

8 gms. of oxygen unite with

$$\frac{0.30}{0.82} \times 8 = 2.93 \text{ gms. of carbon.}$$

Since 8 is the equivalent of oxygen, therefore the equivalent of carbon = 2.93.

EXPT. 278.—To find the Equivalent of Sulphur.—Repeat Expt. 194, in which we found the composition of sulphur dioxide [J] by weight. Here we found that 1.02 gms. of sulphur united with 1.03 gms. of oxygen, so that, since

1.03 gms. of oxygen unite with 1.02 gms. of sulphur;

1 gm. of oxygen unites with $\frac{1.02}{1.03}$ gms. of sulphur,

and 8 gms. of oxygen unite with

$$\frac{1.02}{1.03} \times 8 = 7.92 \text{ gms. of sulphur.}$$

Hence the equivalent of sulphur is 7.92 (accurately = 8).

Now, it is obvious that in the above experiments all we need to know is the weight of the element and the weight of oxide formed, or the weight of oxygen used, so that, provided we begin with a weighed amount of metal and weigh the oxide produced, it is quite immaterial how that oxide is formed.

EXPT. 279.—To determine the equivalent of Tin (Antimony).—Weigh a crucible and lid, and place in it about $\frac{1}{2}$ gm. of pure thin tinfoil, and reweigh. Add strong nitric acid, spot by spot, from the dropping-tube,

keeping the crucible well covered as shown. When no more tin remains, and a little excess of acid has been added, dry gently, and finally ignite strongly. Cool and weigh the tin oxide [L]; add a little more acid, re-heat, cool, and reweigh until of *constant weight*; then calculate as above.



FIG. 91.

EXPT. 280.—**To find the Equivalent of Metals** (*e.g.*, Zinc, Copper, Magnesium, etc.).—Weigh out $\frac{1}{2}$ gm. of zinc, and place it in a weighed porcelain dish with stirring rod. Place on this a cover-glass (Why?), and add a c.cm. or two of strong nitric acid until the metal is dissolved. Evaporate the solution of zinc nitrate to a small bulk, and cautiously wash the liquid from the cover-glass into the dish with a c.cm. or so of distilled water. Carefully evaporate to dryness, stirring continually. (Why?) Ignite the solid zinc nitrate; brown fumes escape (What are they?), and zinc oxide [M] remains. Cool and weigh; reheat, cool and reweigh until of *constant weight*. Calculate as before, and compare your result with the equivalent obtained by other methods.

It would also be quite satisfactory if we began with a weighed amount of the oxide and found the weight of metal in it. Hence we may find equivalents by the **reduction of oxides**, etc.

EXPT. 281.—**To find the Equivalent of Mercury** (Silver, etc.).—Weigh a hard-glass test-tube, with its end loosely plugged with cotton-wool; add about 4 gms. of red mercuric oxide, and reweigh. Now heat the oxide [N] strongly, holding the tube horizontally until all the oxide has disappeared; cool and weigh the tube and residual mercury. Reheat until of *constant weight*, and calculate as before.

For the metals whose oxides are not decomposed by heat the following experiment may be employed.

EXPT. 282.—To find the Equivalent of a Metal (*e.g.*, Copper, Lead, etc.).—Weigh a porcelain boat (*b*); place in it a gram or two of pure dry oxide of lead [*O*], and reweigh. Place this boat and contents in a hard-glass tube (Fig. 27, *a*). Pass dry hydrogen or coal-gas through the apparatus until all air is expelled (Why?); light the issuing gas at *d*, and heat strongly. When the reduction is complete, cool, whilst hydrogen is still passing (Why?), and reweigh when cold. Reheat in hydrogen, cool, and reweigh until of *constant weight*; then calculate as above.

EXPT. 283.—To find the Equivalent of Silver by Reduction of its Chloride.—Weigh a porcelain boat, add about a gram of silver chloride [*A*], and reweigh. Reduce it in hydrogen, as in the preceding experiment, calculating as below:

RESULTS—

Weight of boat and chloride = 4.39 gms.

Constant weight of boat and silver = 4.21 „

∴ Loss = weight of chlorine = 0.18 gm.

Constant weight of boat and silver = 4.21 gms.

Weight of boat alone = 3.68 „

∴ Weight of silver = 0.53 gm.

Now, in Expt. 251 we found the equivalent of chlorine to be 35.5, so that

0.18 gm. of chlorine combines with 0.53 gm. of silver,

1 gm. of chlorine combines with $\frac{0.53}{0.18}$ gm. of silver,

and 35.5 gms. of chlorine combine with

$$\frac{0.53 \times 35.5}{0.18} = 104.5 \text{ gms. of silver.}$$

Since 35.5 is the equivalent of chlorine, therefore the **equivalent of silver = 104.5** (accurately = 108).

Other Methods.

EXPT. 284.—To find the Equivalent of Copper by Replacement with Zinc (Iron).—Weigh out about $\frac{1}{2}$ gm. of zinc foil, and place it in a weighed dish. Cover the foil with a solution of copper sulphate. The zinc dissolves, displacing the copper, which is precipitated [P]. When the action is complete, decant off the residual solution, and wash by decantation with hot water until all the copper sulphate is removed. Dry the copper in a steam oven or at a gentle heat, avoiding oxidation. (What effect would this have on your result?) Cool and weigh, and repeat the washing (Why?) and drying processes until of constant weight. Calculate as below:

RESULTS—

Equivalent of zinc = 32.5 gms.

Weight of zinc = 0.50 gm.

Constant weight of dish and copper = 27.61 gms.

Weight of dish alone = 27.13 „

Weight of copper = 0.48 gm.

So that 0.50 gm. of zinc replaces 0.48 gm. of copper,

1 gm. of zinc replaces $\frac{0.48}{0.50}$ gm. of copper,

and 32.5 gms. of zinc replace

$$\frac{0.48 \times 32.5}{0.50} = 31.8 \text{ gms. of copper.}$$

Thus the equivalent of copper = 31.8.

EXERCISES.—1. Determine the equivalent of iron as above, of silver by replacement from silver nitrate with magnesium, and of lead by its replacement from lead acetate with zinc.

2. Find the equivalent of silver by dissolving $\frac{1}{2}$ gm. of silver in dilute nitric acid, and precipitating it completely as chloride with hydrochloric acid. Wash and dry the silver chloride, and weigh. (Equivalent of chlorine = 35.5.)

3. Find the equivalent of mercury by dissolving 1 gm. of mercury in warm strong nitric acid. Boil off the nitric acid on a water-bath, dilute, add a spot or two of hydrochloric acid, and saturate with hydrogen sulphide. Filter and wash as described; dry and weigh the mercuric sulphide. (Equivalent of sulphur = 16.)

4. Determine the equivalent of copper by dissolving $\frac{1}{2}$ gm. of copper filings in dilute nitric acid, and adding sodium hydrate until strongly alkaline. Boil until quite black, and filter; wash and dry the precipitated copper oxide, and calculate the equivalent.

5. Determine the equivalent of lead by dissolving $\frac{1}{2}$ gm. of lead in dilute nitric acid, and evaporating the solution to dryness on a water-bath. Dissolve the salt in water; cool, and add a solution of potassium iodide until no more precipitate is produced. Filter; wash with cold water; dry and weigh the lead iodide. Calculate the equivalent. (Equivalent of iodine = 127.)

In all the above experiments in some way or another not only are the elements concerned in the changes occurring, but compounds, too, are formed or decomposed, and it will not be surprising to learn that the term "equivalent" has been suitably extended to embrace compounds as well as elements. Moreover, just as equivalents of elements react together, so the *equivalents of elements react with equivalents of compounds*, and the equivalents of the latter with one another. Hence the same unit is employed in their measurement, and some other uses of the term *equivalent* will be discussed, more particularly in their relation to acids, bases, and salts.

The Equivalents of Compounds.

These are all referred to the same unit—viz., the equivalent of hydrogen or of some other element—*e.g.*, the acids. These all contain replaceable hydrogen, and we define the **equivalent of an acid as the amount that contains unit weight of replaceable hydrogen**. Obviously, therefore, it is also the amount which dissolves or reacts with the equivalent of a metal.

EXPT. 285.—Direct Determination of the Equivalent of an Acid.—When we found the equivalent of a metal, we weighed the metal and took an excess of the acid; so for this determination we must take a weighed amount of the acid and excess of a metal. Place in the flask (Fig. 63, *b*) an excess of zinc, and cover it with water. In the small tube (*a*) put 5 gms. of strong sulphuric acid [**Q**]. (Your instructor will tell you its volume, and you can use a burette to measure it.) Fit up the apparatus as in Expt. 152, taking care that the acid does not come in contact with the metal until all is air-tight. When this is so, lower the tube of acid, and allow the action to commence; collect the gas evolved, and find its volume. Work out your results as below:

RESULTS—

Weight of acid = 5.39 gms.

Volume of hydrogen = 1,200 c.cms.

Weight of hydrogen = $\frac{1200 \times 0.09}{1000}$ gms. = 0.108 gm.

Thus, 0.108 gm. of hydrogen contained by 5.39 gms. of acid,

1 gm. of hydrogen contained by

$$\frac{5.39}{0.108} = 49.8 \text{ gms. of acid.}$$

So that the **equivalent of this acid = 49.8**.

For indirect methods of finding equivalent of acids, see works on volumetric analysis.

Bases.—These bodies all neutralize acids to form salts; hence the **equivalent of a base is the amount that will neutralize an equivalent of an acid.**

EXPT. 286.—To find the Equivalent of a Base.—Weigh out about 5 gms. of dry copper oxide, place it in a dish, and cover it with water. Add to this a weighed amount (2 gms.) of sulphuric acid [R], with an equivalent of 49.0, and warm them together. Test the solution with litmus-paper, and, when neutral, decant the whole through a funnel fitted with a double filter-paper, as in Expt. 43. Wash the residue from the dish into the filter and wash the undissolved copper oxide with hot water until the last spot on evaporation leaves no residue. Dry in a steam-oven or drying-cone, and weigh. Calculate as below:

RESULTS—

Weight of acid = 1.98 gms.

Weight of oxide taken = 5.32 gms.

Weight of oxide left = 3.74 „

Weight of oxide dissolved = 1.58 gms.

Now, 1.98 gms. of acid are neutralized by 1.58 gms. of copper oxide,

1 gm. of acid is neutralized by $\frac{1.58}{1.98}$ gms. of copper oxide,

49 gms. of acid are neutralized by $\frac{1.58 \times 49}{1.98} = 39.27$ gms. of copper oxide.

Since the equivalent of the acid is 49, therefore the **equivalent of copper oxide is 39.27.**

For indirect methods of finding the equivalent of bases and of alkalies, a work on volumetric analysis must be consulted.

Salts.—Since salts are formed by neutralizing acids and bases, the **equivalent of a salt** is that amount of salt formed by neutralizing the equivalent of an acid or a base.

EXPT. 287.—To find the Equivalent of a Salt.—In the previous experiment the filtrate which contains the salt is evaporated to dryness (Why ?) in a weighed dish and weighed. It is reheated until of constant weight, and again weighed.

RESULTS—

Weight of acid = 1.98 gms.

Constant weight of dish and copper

sulphate = 29.74 gms.

Weight of dish alone = 26.54 „

Weight of copper sulphate = 3.20 gms.

Thus, 1.98 gms. of acid form 3.20 gms. of the salt,

1 gm. of acid forms $\frac{3.20}{1.98}$ gms. of the salt,

and 49 gms. of acid form $\frac{3.20 \times 49}{1.98} = 78.8$ gms. of the salt.

So that the **equivalent of copper sulphate** is 78.8.

EXPT. 288.—To determine the Equivalent of Sodium Chloride.—Place in a beaker containing 25 c.cms. of water a weighed amount of strong hydrochloric acid. Neutralize this solution with caustic soda solution [S] as in Expt. 170. When neutral, evaporate the liquid to dryness, rinsing out the beaker several times with small quantities of water, in a weighed dish with a rod; cool, and reweigh. Reheat, cool and reweigh; then calculate as above.

EXERCISES.—1. Find the equivalent of marble by placing a weighed piece in 25 c.cms. of water in a beaker and adding a weighed quantity of strong hydrochloric

or nitric acid. When action over, wash and dry the marble, and reweigh.

2. Find the equivalent of zinc carbonate by placing a weighed excess of the solid (10 gms.) into 25 c.cms. of water in a beaker containing 5 gms. of sulphuric acid. When the action is over, filter, wash and dry the undissolved solid, and weigh. Repeat this for other carbonates.

3. Find the equivalent of lead oxide, iron oxide, zinc oxide, etc., as in Expt. 286, and of lead nitrate, iron sulphate, zinc chloride.

4. Find the equivalent of potassium nitrate, potassium sulphate, ammonium chloride, etc., as in Expt. 288.

For further details and other uses of the term "equivalent," the student is referred to more advanced textbooks.

REVISION QUESTIONS.

CHAPTERS XVIII. AND XIX.

1. Mention the chief uses of sulphuric acid, and give experiments to illustrate your answer.
2. Sulphur is a non-metal. What evidence would you bring forward in favour of this statement ?
3. How is sulphuric acid made ? What chemical changes take place when sulphuric acid reacts with (a) calcium carbonate, (b) lead peroxide, (c) sodium hydroxide ?
[C.L.J., 1909.]
4. How is sulphur dioxide prepared ? What chemical changes take place when sulphur dioxide is passed into (a) chlorine water, (b) nitric acid ?
[C.L.J., 1909.]
5. How would you prove experimentally that the weight of the product obtained when sulphur is burned in oxygen is greater than the weight of sulphur that has disappeared ? Draw a diagram of the apparatus you would use. How may the gaseous product be regained after this experiment ? How would you identify it ?
[L.M., 1911.]
6. State the chemical properties of iron pyrites. How could

you obtain from this (1) some sulphur, (2) some sulphur dioxide, (3) some hydrogen sulphide ?

7. Give an account of the various kinds of sulphur, and describe the changes that occur when sulphur is gradually heated in a test-tube to a high temperature. [C.L.J., 1911.
8. If you were supplied with some sulphur, how would you prepare an acid ?

CHAPTERS XX. AND XXI.

1. How would you prepare nitric oxide gas ? Give an account of its properties. [C.L.J., 1907.
2. How would you prepare and collect nitrous oxide ? What experiments would you yourself propose to make to show that this gas contains its own volume of nitrogen [L.M., 1912.
3. How is nitric acid prepared ? State its properties. What are the chief chemical properties of this acid ? [C.L.J., 1907.
4. How would you obtain from nitric acid a sample of nitrous oxide and nitrogen peroxide ? How are nitrates formed in the soil ?
5. Nitric acid and nitrates are said to be oxidizing agents. Give experiments to illustrate this.
6. What experiments would you use to determine whether a given substance was a nitrate or not ?
7. Describe the preparation of nitric acid. How does dilute nitric acid act upon (a) copper, (b) caustic soda, (c) calcium carbonate ? [C.L.J., 1911.

CHAPTER XXII.

1. How would you prepare hydrogen chloride (gas) from some common salt ? Draw a neat diagram of the apparatus you would use, mentioning the use of each part.
2. What experiments would you do in order to distinguish between nitric, hydrochloric, and sulphuric acid ?
3. Describe with a sketch of the apparatus how you would prepare a solution of hydrochloric acid. Given a supply of Iceland spar—i.e., calcium carbonate—suggest a method by which you could employ this substance to find the percentage by weight of hydrogen chloride in your solution. [L.M., 1911.

4. How is chlorine prepared? Describe experiments you would perform with this gas to illustrate its character.
[C.L.J., 1907.]
5. Name three substances occurring commonly in nature which contain chlorine. Explain how you would prepare from one of these (1) chlorine, (2) hydrogen chloride.
6. "Hydrogen is said to have a great affinity for chlorine." Describe experiments which verify this statement.
7. Chlorine is a vigorous supporter of combustion. Give experiments to justify this statement.
8. Explain, giving experiments, the bleaching action of chlorine.
9. What are chlorides? How are they produced? Describe one experiment in which a metallic chloride and one in which a non-metallic chloride is obtained, and compare the properties of the products.
10. Hydrogen chloride is composed of equal volumes of hydrogen and chlorine. What experiments have you seen which support this statement?

CHAPTER XXIII.

1. Describe an experiment to show that when a candle burns the matter of which it consists is not destroyed.
[C.L.J., 1911.]
2. Three samples of lead oxide prepared in different ways were reduced by heating them in a stream of hydrogen. The results of the experiments were—(a) 6.36 gms. of oxide produced 5.90 gms. of lead; (b) 4.96 gms. of oxide produced 4.60 gms. of lead; (c) 4.02 gms. of oxide produced 3.73 gms. of lead. Are these results in accordance with the law of definite proportion?
3. What do you understand by a "law"? Mention two chemical laws, and explain how they have been arrived at.
4. Describe any two simple experiments which may be carried out in the laboratory which illustrate the following statement: "However a chemical compound is prepared, it is always found to be composed of the same elements combined in the same proportions."

CHAPTER XXIV.

N.B.—1,000 c.cms. of hydrogen weigh 0.09 gm.

1. What do you mean by the term "equivalent weight" ?
In an experiment 0.3 gm. of a metal was found to liberate 85 c.cms. of hydrogen at 0° C. and 760 mm. Find the equivalent of the metal. [C.L.J., 1911.]
2. Copper oxide was reduced in hydrogen, and the water produced was collected in a U-tube of calcium chloride. Draw a sketch of the apparatus. The following results were obtained: Weight of copper oxide, 15.43 gms.; weight of copper, 12.30 gms.; weight of water formed, 3.52 gms. From these calculate (1) the equivalent of oxygen; (2) the equivalent of copper.
3. Silver chloride reduced in coal-gas lost 7.09 gms., whilst the hydrogen chloride produced weighed 7.29 gms. Find the equivalent of chlorine.
4. If 0.10 gm. of a metal when dissolved in dilute sulphuric acid gives off 34.15 c.cms. of hydrogen, what is its equivalent ?
5. If 0.50 gm. of zinc dissolved in dilute hydrochloric acid yields 170 c.cms. of hydrogen, and 0.30 gm. of zinc warmed with caustic soda solution gives off 102 c.cms. of the gas, calculate the equivalent of zinc from each of these results. What does this prove ?
6. 0.20 gm. of magnesium reacts with dilute acid, giving off 186 c.cms. of hydrogen. Calculate its equivalent.
7. 0.20 gm. of sodium acting on water sets free 97.5 c.cms. of hydrogen. What is its equivalent ? Draw a sketch of the apparatus you would use.
8. 0.30 gm. of aluminium liberates 373 c.cms. of hydrogen when warmed with caustic soda solution. Find its equivalent.
9. 0.20 gm. of potassium was dissolved in mercury and this amalgam put in water; 57.4 c.cms. of hydrogen was liberated, and mercury left. Find the equivalent of potassium ?
10. 0.25 gm. of sodium was first dissolved in mercury, and this was placed in dilute sulphuric acid. Mercury remained, and 121.9 c.cms. of hydrogen were evolved. Find the equivalent of sodium.

11. [N.B.—The equivalent of oxygen is 8.] Calculate the equivalent of copper if 73.34 gms. of the metal form 91.79 gms. of the oxide.
12. When steam was passed over 0.5 gm. of heated magnesium, 0.83 gm. of magnesia was formed. What is the equivalent of magnesium ?
13. 0.5 gm. of sulphur burned in oxygen yielded 0.99 gm. of sulphur dioxide. Find the equivalent of sulphur.
14. 1 gm. of copper heated in air until constant in weight then weighed 1.26 gms. What was produced, and what is the equivalent of copper ?
15. $\frac{1}{2}$ gm. of iron filings heated in steam yielded 0.69 gm. of residue. What is the equivalent of iron ?
16. $\frac{1}{2}$ gm. of pure sugar charcoal burned in oxygen yielded 1.83 gms. of carbonic acid gas. Give full instructions for carrying out this experiment, and find the equivalent of carbon.
17. 1.5 gms. of lead heated in oxygen formed 1.62 gms. of lead oxide. What is the equivalent of lead ?
18. When $\frac{1}{2}$ gm. of phosphorus was heated in oxygen, the oxide formed weighed 1.15 gms. What is the equivalent of phosphorus ?
19. 0.3 gm. of magnesium was dissolved in nitric acid, and the solution was evaporated, and the residue strongly ignited until no more brown fumes escaped. The residue weighed 0.5 gm. What was the residue, and what is the equivalent of magnesium ?
20. Nitric oxide was reduced by red-hot copper. The nitrogen collected had a volume of 1,064 c.cms., and the gain in weight of the copper was 1.52 gms. Calculate the equivalent of nitrogen if its density is 0.00126 gm. per c.c.
21. 2.5 gms. of lead oxide when reduced in coal-gas left 2.32 gms. of lead. What of the equivalent of lead ?
22. 3.04 gms. of a metallic oxide, when reduced in hydrogen, left 2.64 gms. of the metal. Find the equivalent of the metal.
23. 0.70 gm. of a metal produced 0.86 gm. of oxide. What is the equivalent of the metal ?
24. A metal *A* produced when oxidized 1.25 gms. of its oxide, and an equal weight of *B* produced 1.21 gms. of oxide. If the equivalent of *B* is 36, what is the equivalent of *A* ?

25. Equal weights of two metals, X and Y , liberated 68 c.cms. and 55 c.cms. of hydrogen respectively. If the equivalent of X is 39, find the equivalent of Y .
26. 1.45 gms. of a metallic oxide were completely reduced in hydrogen, and 0.59 gm. of water was formed. Assuming that 9 gms. of water contain 1 gm. of hydrogen, calculate the equivalent of the metal.
27. 1.071 gms. of tin dissolved in strong hydrochloric acid liberated 200 c.cms. of hydrogen; 1.661 gms. of an oxide of tin left 1.309 gms. of the metal on reduction. Find the equivalent of tin from each of these experiments, and compare your results. Explain, if you can, the results obtained.
28. Laughing-gas reduced by red-hot copper gave the following results: (a) Volume of nitrogen = 2.016 c.cms.; (b) gain in weight of copper = 1.44 gms. If the density of nitrogen = 0.00126 gm. per c.cm., find (1) the equivalent of nitrogen; (2) the percentage composition of laughing-gas.
29. It was found that 2.5 gms. of lead monoxide (litharge) left 2.32 gms. of lead when reduced in hydrogen. 3.20 gms. of red lead similarly treated left 2.90 gms. of lead. Calculate the equivalent of lead from these results, and explain the results obtained.
30. x gms. of a metal were dissolved in acid, and the hydrogen obtained was burned to form x gms. of water. What is the equivalent of the metal?
31. 0.5 gm. of zinc is dissolved in acid, and the liberated hydrogen burned. If the equivalent of zinc is 32.5, what weight of water is obtained?
32. A certain quantity of a metal when dissolved in acid yielded hydrogen, which, on burning, produced three-quarter times the weight of water as of metal taken. What is the equivalent of the metal?
33. Hydrogen was passed over excess of copper oxide, and the oxide lost 59.78 gms., whilst the water formed weighed 67.282 gms. Calculate the equivalent of oxygen if hydrogen = 1.
34. When 2.5 gms. of silver were heated in chlorine, 3.32 gms. of silver chloride were formed. If the equivalent of chlorine = 35.5, find the equivalent of silver.
35. 0.4122 gm. of a metal, when heated in chlorine until the

metal would combine with no more gas, the weight of the product was found to be 0.6249 gm. Draw the apparatus you would use, and calculate the equivalent of the metal if the equivalent of chlorine is 35.5.

36. 5 gms. of mercuric oxide, when heated, left 4.63 gms. of mercury. What is the equivalent of mercury if that of oxygen is 8?
37. 1.59 gms. of silver chloride reduced in hydrogen left 1.20 gms. of silver, and produced 0.401 gm. of hydrogen chloride. Calculate the equivalent (1) of chlorine, (2) of silver.
38. 0.5 gm. of sodium heated in chlorine formed 1.27 gms. of common salt. If the equivalent of chlorine = 35.5, find the equivalent of sodium.
39. When 4.64 gms. of silver oxide were heated, 4.32 gms. of silver remained. What is the equivalent of silver?
40. 1.3 gms. of zinc dissolved in copper sulphate solution precipitated 1.26 gms. of copper. If the equivalent of zinc is 32.5, find the equivalent of copper.
41. 3.52 gms. of copper are precipitated from a solution of a copper salt by 3.12 gms. of iron. If the equivalent of copper is 31.5, find the equivalent of iron.
42. 0.4 gm. of magnesium ribbon immersed in silver nitrate solution deposits 3.6 gms. of silver. Determine the equivalent of silver if that of magnesium is 12.
43. When 0.26 gm. of zinc dissolve in lead acetate solution, 0.82 gm. of lead is precipitated. What is the equivalent of lead if that of zinc is 32.5?
44. 1 gm. of zinc displaces from a solution of a metallic salt 0.96 gm. of the metal. Calculate the equivalent of the metal if that of zinc is 32.5.
45. Equal weights of zinc put into the solutions of the sulphates of two metals, *A* and *B*, set free 0.53 gm. of *A* and 0.94 gm. of *B*. The equivalent of *B* is 64; find the equivalent of *A*.
46. 0.5 gm. of zinc liberated 172 c.cms. of hydrogen when placed in an acid. A second 0.5 gm. of zinc placed in the acid containing some gold chloride set free 42 c.cms. of hydrogen and 0.73 gm. of gold. Find the equivalent of gold.
47. 2.5 gms. of caustic soda were exactly neutralized by hydrochloric acid, and the solution carefully evapor-

ated. The dry common salt weighed 3.66 gms. Find the equivalent of common salt if that of caustic soda = 40.

48. 1.5 gms. of sulphuric acid when diluted and treated with excess of zinc set free 343 c.cms. of hydrogen. What is the equivalent of the acid ?
49. 5 gms. of sulphuric acid were diluted and warmed with 10 gms. of copper oxide. The undissolved copper oxide, when washed and dried, weighed 5.97 gms. If the equivalent of the acid is 49, find the equivalent of copper oxide.
50. 5 gms. of hydrochloric acid gas dissolved in water and treated with excess of magnesium gave off 1,525 c.cms. of hydrogen. Find the equivalent of the acid.
51. 2.5 gms. of sulphuric acid required 2.04 gms. of caustic soda to just neutralize it, and the dry salt obtained by evaporation weighed 3.62 gms. Given the equivalent of the acid as 49, find the equivalent (1) of caustic soda, (2) of the salt.
52. 7.3 gms. of hydrochloric acid were just neutralized by 28.6 gms. of washing-soda. If the equivalent of the acid is 36.5, find the equivalent of washing-soda.
53. When a piece of marble is immersed in a solution containing 5 gms. of hydrogen chloride, it is found to lose in weight 6.85 gms. The equivalent of hydrogen chloride being 36.5, find the equivalent of marble.
54. 1.5 gms. of marble, when heated, lose 0.66 gm. If the equivalent of marble is 50, find the equivalent (1) of lime, (2) of carbon dioxide.

REVISION TESTS.

L.

1. Give an account of the gas produced when sulphur is burned in the air, and explain how this gas may be converted into sulphuric acid. Mention some of the uses of sulphuric acid in chemistry. [C.W.B.J., 1907.
2. What experiments would you make in support of the statement that nitric acid is an oxidizing agent? Starting from nitric acid, how would you obtain specimens of (a) oxygen, (b) laughing-gas? [C.W.B.S., 1908.
3. How is hydrogen chloride prepared for use in the laboratory? Why do we regard hydrogen chloride as a compound substance, and consider its constituents hydrogen and chlorine to be elements? Quote any experiments you can think of that will help to make this clear. [C.W.B.J., 1906.
4. What is meant by "combustion"? Describe two experiments showing that combustion can take place in gases other than oxygen, and write brief accounts of the products of such combustions. [C.W.B.J., 1909.
5. Aluminium foil is soluble in warm caustic soda solution, hydrogen being evolved. Explain clearly any method by which you would propose to utilize this fact in order to determine the equivalent of the metal. [C.W.B.J., 1908.

M.

1. Describe the gases that are produced when sulphuric acid acts on zinc and copper. Say how you would prepare these gases from the above materials, and how you would obtain crystals of the sulphates of these two metals. [C.W.B.J., 1906.
2. How is nitric acid prepared on a small scale in the laboratory? Mention a few instances of its oxidizing action, and explain how you would distinguish it from sulphuric acid and from a strong solution of hydrochloric acid. [C.W.B.J., 1908

3. What is chlorine ? How is it prepared from hydrochloric acid ? Describe clearly how you would obtain a few jars of this gas, and study its bleaching action.
[C.W.B.J., 1907.]
4. What evidence have we that every piece of marble is composed of lime and carbon dioxide in fixed proportions ? Describe some experiments to illustrate your answer.
[C.W.B.J., 1907.]

N.

1. Write a short account of the chemical properties of sulphur and describe in detail all that you see when you heat a piece of sulphur in a test-tube gradually to its boiling-point, and then allow it to cool again. [C.W.B.J., 1906.]
2. Describe in some detail how you would investigate the action of copper on nitric acid, showing how you would isolate specimens of any substances other than water formed during the action. Describe, as far as you can, the properties which the above products possess.
[C.W.B.J., 1910.]
3. How would you prepare and collect gaseous hydrochloric acid, starting from common salt ? Describe the properties of the substances formed when (a) lime, (b) ammonia, are dissolved in a solution of hydrogen chloride.
[C.W.B.J., 1909.]
4. Explain clearly the use in chemistry of the term "equivalent weight" (a) when applied to an element, such as oxygen or zinc, (b) when applied to an acid. Describe an experiment by which you could find the equivalent of copper. If 0.245 gm. of a certain acid, when treated with excess of zinc, gives off 55.5 c.cms. of hydrogen, what is the equivalent of the acid ? [1 litre of hydrogen weighs 0.09 gm.]
[C.W.B.J., 1907.]
5. Give an account of any experiments you have made to illustrate the law of definite proportions.
[C.W.B.S., 1908.]

O.

1. Show from the properties of sulphur that an element may exist in more than one form. [C.W.B.J., 1907.]
2. Describe fully the preparation from salt of a jar full of chlorine. Explain, with two examples, why chlorine is regarded as an oxidizing agent. [L.M., 1912.]
3. Elements behave in various ways towards water. Some merely dissolve in it; others act upon it either when the water is liquid and at ordinary temperature or at temperatures above 100° C. Give an account of the behaviour in this respect of (a) sodium, (b) carbon, (c) oxygen, (d) chlorine, (e) magnesium, (f) iron. [L.M., 1912.]
4. If you were given some ammonium nitrate, how would you prepare from it and collect (a) an oxide of nitrogen, (b) ammonia, (c) nitrogen? Give equations and draw diagrams of the apparatus you would use. [L.M., 1912.]
5. What is meant by the equivalent weight of an element? Give full instructions for determining the equivalent of some one metal. A certain quantity of a metal is dissolved in an acid. The evolved hydrogen, when burned in air, produces a weight of water exactly equal to the weight of the metal dissolved. What is the equivalent of the metal? [C.W.B.J., 1909.]
6. Describe, with careful attention to detail, the *chemical* methods you would use to obtain each of the metals from a mixture of finely divided copper and iron. [L.M., 1912.]

P.

1. Describe the various modifications of sulphur, and explain exactly how you would show that each is composed wholly of the same elementary substance. [C.W.B.S., 1910.]
2. Describe the preparation of a small quantity of a strong solution of hydrogen chloride, starting from common salt. How would you show that this gas is a compound, and not merely a mixture of hydrogen and chlorine? [C.W.B.S., 1909.]

3. Describe the preparation of any oxide of nitrogen, starting from nitric acid. Explain fully how you would obtain a sample of nitrogen from the oxide you select.
[C.W.B.S., 1910.]
4. Write an account of the action of heat on various nitrates, giving sketches to explain how you would collect a pure specimen of each of the gaseous substances you mention.
[C.W.B.S., 1909.]
5. Describe the chemical changes which take place when (a) sodium, (b) sulphur, (c) sulphuretted hydrogen, (d) iron, are burned in oxygen.
[C.L.J., 1907.]
6. How would you prepare a small quantity of sulphuric acid, starting with sulphur? Describe experiments to illustrate (a) its strong affinity for water, (b) its action on common metals. In the preparation of nitric acid why would you use sulphuric acid rather than hydrochloric acid?
[L.M., 1912.]

PRACTICAL TESTS.

1. Determine the acid in the substance *O* supplied.
2. Identify the simple salts *P*.
3. Describe what you observe when the compound *Q* is heated. Examine and identify the products. C.L.J., 1909.
4. Identify the compound *R*.
5. Determine chemically whether *S* dissolves in dilute acids and alkalies. Identify any gas evolved in each case.
6. You are supplied with the oxide of a metal *T*. Prove experimentally that it is an oxide, and find the metal present.
7. You are given some pieces of a metal, *U*. Verify one important law by experiments on this.
8. Find the weight of the metal *V* that will unite with 16 gms. of oxygen.
9. Heat a little of *W* in a crucible with a lid, and maintain it at a low red heat for three minutes. When it is cold, act upon the residue with dilute hydrochloric acid, and identify the gas evolved. *Prove your conclusion by confirmatory experiments.* Filter the solution obtained, and detect any metal contained in it. [C.L.J., 1911.]

EXAMINATION PAPERS.

CENTRAL WELSH BOARD (JUNIOR) CHEMISTRY.

(Five only to be attempted.) 2½ Hours.

1911.

1. Some sand and saltpetre are placed in a large bottle, which is afterwards half filled with freshly boiled distilled water. The bottle is then corked and vigorously shaken; the cork is removed and replaced, and the bottle again shaken. Show how you would recover the sand and saltpetre, and how you would endeavour to find out whether any air has dissolved in the liquid.
2. Name and describe three substances which evolve oxygen when strongly heated in hard-glass test-tubes. Describe also, as far as you can, the appearance and chemical properties of any substances which remain in the test-tubes after the heating.
3. The examination syllabus contains the words, "Reduction of the oxides of iron, lead, and copper by carbon and hydrogen." Show how you would reduce a metallic oxide (a) by carbon, (b) by hydrogen, briefly explaining how process (b) might be used to determine the equivalent of the metal.
4. Sketch and describe an apparatus suitable for preparing a jar of sulphur dioxide. By what experiments would you distinguish between a jar of sulphur dioxide and one of carbon dioxide?
5. How would you ascertain the weight of gas which would be expelled from a kilogramme of chalk on treatment with excess of dilute acid? Carefully note any precautions which should be observed during the experiment.
6. What are *salts*? Describe three methods of producing them, giving one illustration of each method.
7. Write a short essay upon natural waters. Explain the causes which produce hardness in many samples of natural water.

1912.

1. Describe, with sketches, the experiments you would make to show that (a) sodium, (b) magnesium, reacts with water. Write brief accounts of all the substances produced in each case.
2. How could you prove that your own breath contains carbon dioxide? How could you separate and collect a small quantity of the dioxide from your breath? Give sketches of the apparatus you would use.
3. Give an account of the experiments you would undertake to show that air contains an active gas, identical in properties with that obtained on heating chlorate of potassium.
4. Clearly explain the difference between *acidic* and *basic* oxides. Briefly describe the preparation of one member of each class.
5. What happens when strong sulphuric acid is heated (a) with nitre, (b) with common salt? Sketch the apparatus with which these reactions are usually carried out, and describe briefly how the products may be reconverted into nitre and common salt respectively.
6. Mention, *without detail*, any three methods by which the equivalent of metals has been determined. Describe *in detail* how you would determine the equivalent of copper.
7. Explain carefully how you would ascertain the total weight of hydrogen displaceable from dilute sulphuric acid by a kilogramme of zinc.

1913.

1. What do you understand by the terms "element" and "compound"? What experiments would you perform with (a) marble, (b) water, (c) copper, in order to determine whether the substance is an element or a compound?
2. How would you proceed to convert magnesium and sulphur into their oxides? In the former case what observations are necessary in order to determine the equivalent of the substance?

3. Give a list of the various gases which enter into the composition of the atmosphere, and describe the most accurate method known to you for determining the percentage amount of one of the constituents present. By what agencies is the comparative constancy of composition of the air maintained ?
4. Explain how you would proceed to determine the volume of gas evolved by heating 100 gms. of mercuric oxide. Mention all essential precautions, and draw a diagram of the apparatus.
5. What is caustic soda, and how can it be obtained from washing-soda? Describe the action of dilute hydrochloric acid on caustic soda and washing-soda respectively.
6. You are required to investigate the purity of your laboratory water-supply. What impurities are likely to be present, and how may they be detected? What means are applicable for removing these impurities ?
7. What do you understand by "allotropy" ? Describe the preparation of the various allotropic forms of sulphur.

****PART IV**

CHAPTER XXV

OTHER CHEMICAL LAWS AND THE ATOMIC THEORY

AN examination of the values of the equivalents found in Chapter XXIII. reveals a somewhat curious fact. While some elements, like chlorine, zinc, etc., have always the same combining weight, other elements, *e.g.*, sulphur, lead, nitrogen, iron, etc., have more than one value. Sulphur has *two* combining weights, lead *three*, nitrogen *five*, and so on—*i.e.*, **the equivalent varies**. A closer investigation of these different values shows us that a *simple relation exists between the equivalent values* for the same element. For instance, sulphur has the values 16 and 8—*i.e.*, in the ratio of 2 to 1; lead with the values 103.5, 77.6, and 51.75, in the ratio of 4 to 3 to 2, and so on. The fact that **these equivalents are in a simple multiple ratio** are embodied in a law—**THE LAW OF MULTIPLE PROPORTIONS**—which states: **When two elements unite in more than one proportion by weight for a fixed weight of one element, there is a simple multiple ratio between the weights of the other element.**

Let us consider this in detail, and take for our experiments the two oxides of copper. The first,

black cupric oxide, *always has the same composition by weight*; while the second, red cuprous oxide, *always has the same composition by weight—i.e., they each obey the law of definite proportion (or constant composition).*

EXPT. 289.—To show that “Copper Oxides” obey the Law of Multiple Proportions.—Repeat Expt. 282, using *black* copper oxide. From this we find that black copper oxide contains 79.75 per cent. of copper, and 20.25 per cent. of oxygen.

Repeat the experiment, using *red* cuprous oxide, calculating as before. Now we find that

Red cuprous oxide contains 88.73 per cent. of copper, and 11.27 per cent. of oxygen.

Let us **fix** the weight of oxygen as 1. Then this weight of oxygen in

- (1) Black copper oxide unites with $\frac{79.75}{20.25}$; i.e., 3.94 parts of copper.
- (2) Red cuprous oxide unites with $\frac{88.73}{11.27}$; i.e., 7.87 parts of copper.

Hence the ratio of copper which unites with 1 of oxygen in the two compounds is 3.94 : 7.87, or 1 : 2—a **simple multiple ratio**.

EXERCISE.—Verify the law (1) for the three lead oxides as above; (2) for the three oxides of nitrogen as in Expt. 221; (3) for the oxides of iron; (4) for the oxides of tin.

These simple relations in which elements combine together were first clearly perceived by the English scientist, **John Dalton**, c. 1805-1808, and it is to him that we owe a **theory** (*an explanation, or a mental picture*) of the formation of compounds from elements according to these three experimental laws.

Dalton's Atomic Theory.—According to Dalton, **matter consists of small, indivisible particles—the “atoms”**—*each having its own properties, weight, etc. The atoms of the same substance are all exactly alike, but the atoms of one substance are quite different from the atoms of all other substances.*

It must be clearly understood that nothing is implied as to shape, size, etc., of atoms. Whether they are spherical or square, etc., large or small, soft or hard, is not known, and is quite unimportant; so that, although we use objects to illustrate this idea, it is not intended that they should represent atoms except in name only.

In order that atoms shall obey the law of conservation of matter, *atoms enter into chemical actions entirely—i.e., without change of weight.* How does this agree with the other laws of chemical union? Let us represent two kinds of atoms by a heap of shillings for one kind and a heap of pennies for the other. First of all we may make a number of groups, each containing one shilling and one penny. These groups are obviously exactly alike—*i.e., the same atoms in the same proportion; i.e., they obey the law of definite proportion.* We may make a second set of groups, each containing two shillings and one penny. These, again, are obviously all alike—*i.e., obey the law of constant composition—but if we compare them with the first set, we find (1) they are quite different from these—i.e., are different compounds; (2) the number of shillings associated with one penny in the first set is one, and in the second is two—i.e., these different compounds obey the law of multiple proportion.*

Thus the atomic theory, as it is termed, is *consistent* with the laws on which it is based, and is useful. The question may be asked, *Is it true?* The answer is,

We do not know, nor do we greatly care. The theory is a *tool*, like the carpenter's hammer or chisel, and so long as it is satisfactory it is used. When it ceases to be *useful*, it is sharpened up if it is worth it; if not, it is thrown aside and a better one formed. The *justification* for the important place the chemist gives to the *atomic theory* is well founded. The facts it had to explain at its birth were very few, but it explains equally well and readily the multitude of facts of the present day. It has been useful, too, in guiding the chemist to new researches, with uniformly good results.

It is worth noting that the two important laws given above may be expressed in one law of wider application—the **law of equivalents**—viz., **chemical action occurs in quantities proportional to the equivalents or simple multiples of the equivalents.**

So far we have only concerned ourselves with the way in which elements react by weight, but if we reconsider our earlier experiments more carefully, we shall notice that elements seem to react by **volume** in a similar regular manner. We found in Expts. 56 and 89 that water was always composed of two volumes of hydrogen and one volume of oxygen; Expt. 130 taught us that carbon dioxide contained its own volume of oxygen; Expt. 188 that carbon monoxide contained half its volume of oxygen; while similar results were obtained in Expts. 195, 215, 235, etc. *Is there any law connecting the volumes in which gaseous substances react?* Similar experiments to the above were made by the French chemist Gay Lussac, and, as in those mentioned, the experiments were carried out under conditions which remained the same—i.e., the gases were measured under the same temperature and

pressure. His conclusions he stated in the form of a law—viz.:

(Gay Lussac's) Law of Gaseous Volumes (1808).—Gases unite in volumes which bear a simple relation to one another and to the volume of the product, if gaseous.

At first Dalton welcomed this law as confirming his atomic theory, and put forward the *assumption* that "equal volumes of all gases contained the same number of *atoms*." When he applied this assumption to actual facts—*e.g.*, the decomposition of nitrous oxide, in which two volumes of nitrous oxide yielded two volumes of nitrogen and one volume of oxygen—*i.e.*, according to his assumption, two atoms yield two atoms *plus* one atom, or three atoms—he found it led him to the conclusion that atoms *were divisible*. He rejected the assumption, and discredited the **experimental law**. This naturally led many eminent scientists to disregard the atomic theory entirely, and it remained for **Avogadro**, an Italian physicist, in 1811 to reconcile these experimental laws and to set up the atomic theory more firmly than ever in a widened form. He pointed out that Dalton spoke of the atoms of a *substance*, whether that substance was an element or a compound. The latter were obviously divisible by chemical means, and he suggested that *Dalton's atoms* were the *physical final particles* for which he proposed the name **molecule**, and that these could be divided by chemical means into the still smaller *chemical final particles*—the **atoms**. Hence we get from Avogadro (1) the definition of a molecule as the smallest particle of any substance, whether element or compound, which can exist free and retain the properties of the original body; (2) the definition of an atom as the smallest particle of an

element that can take part in any chemical change; and (3) Avogadro's hypothesis—viz., all gases under the same conditions of temperature and pressure contain the same number of molecules.

Using Avogadro's idea, we now see that all discord between Gay Lussac's law of gaseous volumes and the atomic theory vanishes—nay, more, they lend each other mutual support, and strengthen each other. We have seen in Expt. 258 evidence that gases at the moment of liberation in the so-called *nascent* state are particularly active, and we *believe*, therefore, that, under ordinary conditions, atoms do not exist free, but associate together to form molecules. If the atoms within a molecule are *all alike*, it is an *elementary molecule*, whereas if the atoms are *dissimilar*, we have a *compound molecule*.

The Atomic Theory.

Thus, the modern atomic theory is Dalton's theory, as modified by Avogadro. According to this, *matter is supposed to consist of numerous small particles, called "molecules."* These molecules themselves are usually composed of still smaller indivisible particles—the atoms of elements. The atoms of one element are all exactly alike in weight, properties, etc., but totally different from the atoms of all other elements. These atoms enter into chemical actions, according to the experimental laws of chemical combination, without change, in virtue of the action of a force, "*chemical attraction*," or "*chemical affinity*."

Now, although we cannot actually get separate molecules to weigh and measure, we may, by *assuming* Avogadro's hypothesis, obtain their *relative weights* by

weighing equal numbers of molecules or equal volumes. As a rule these are compared with hydrogen as a standard. (Why ?) Now, hydrogen is also ranked as the standard for atomic weights, and it is therefore of the greatest importance to find the **number of atoms in a hydrogen molecule**, or confusion will result. This is determined thus:

Experiment shows that hydrogen unites with chlorine to form hydrogen chloride in the proportion of 1 *volume* of hydrogen to 1 *volume* of chlorine, producing 2 *volumes* of hydrogen chloride, which by *Avogadro's hypothesis* becomes—

n molecules of hydrogen to *n molecules* of chlorine, forming *2n molecules* of hydrogen chloride, or 1 *molecule* hydrogen to 1 *molecule* of chlorine, forming 2 *molecules* of hydrogen chloride.

Now, each molecule of hydrogen chloride must contain *at least* 1 atom of hydrogen, and since 1 molecule of hydrogen is distributed through 2 molecules of hydrogen chloride, therefore 1 *molecule of hydrogen must contain at least 2 atoms* of hydrogen.

Now, *experiment* shows that whatever means we employ, the hydrogen in hydrogen chloride can only be expelled in *one* stage, not, as in water or sulphuric acid, in two stages. Hence, 1 molecule of hydrogen chloride contains *only 1 atom* of hydrogen, and it follows that **1 molecule of hydrogen contains 2 atoms of hydrogen.**

It is obvious that, fixing the **atomic weight of hydrogen as 1**—i.e., the unit or standard of weight—the **molecular weight of hydrogen is 2.** Further, since the

Molecular weight

$$= \frac{\text{weight of 1 molecule of the substance}}{\text{weight of 1 atom of hydrogen}} \quad (A)$$

and the **relative density of a gas**

$$\begin{aligned}
 &= \frac{\text{weight of a gas or vapour}}{\text{weight of an equal volume of hydrogen}} \\
 &= \frac{\text{weight of } n \text{ molecules of a gas}}{\text{weight of } n \text{ molecules of hydrogen}} \\
 &= \frac{\text{weight of 1 molecule of a gas}}{\text{weight of 1 molecule of hydrogen}} \\
 &= \frac{\text{weight of 1 molecule of a gas}}{\text{weight of 2 atoms of hydrogen}} \quad (B)
 \end{aligned}$$

We see, by comparing *A* and *B*, that the **molecular weight of a gas** = $2 \times$ vapour density, or **Mol. Wt.** = $2 \times V.D.$

CHAPTER XXVI

SYMBOLS, FORMULÆ, AND EQUATIONS

FOR convenience, the chemist represents the elementary atoms by **symbols**—*i.e.*, a letter or letters—*representing one atom of that element*, and hence it also represents a definite weight (termed the **atomic weight**), and if gaseous, a definite volume (the **atomic volume**)—*e.g.*, H represents (1) hydrogen (2) 1 atom of hydrogen, (3) 1 part by weight of hydrogen, (4) 1 part by volume of hydrogen. S represents (1) sulphur, (2) 1 atom of sulphur, (3) 32 parts by weight of sulphur, and so on.

A list of the elements, with their symbols and approximate atomic weights, is given on p. 316.

In order to represent *compounds*, **formulæ** are employed, which are merely *collections of symbols which indicate the elements present in the molecule*—*e.g.*, CuO. In order to avoid the repetition of symbols, small foot figures are used to denote the number of atoms of each element present—*e.g.*, H_2SO_4 , instead of HHSOOOO. Now, because symbols are *quantitative*, **formulæ must also be quantitative**, and thus a *formula denotes* (1) the name of the substance; (2) the number and nature of the elementary atoms in a molecule; (3) a definite weight (the **molecular weight** which is the sum of the atomic weights of the elements present): and (4) if gaseous, a definite volume—the **molecular volume**.

TABLE OF ELEMENTS.

Name.	Symbol.	Atomic Weight.	Name.	Symbol.	Atomic Weight.
Aluminium ..	Al	27	Nickel ..	Ni	59
Antimony ..	Sb	120	Niobium ..	Nb	94
Arsenic ..	As	75	Nitrogen ..	N	14
Barium ..	Ba	137	Osmium ..	Os	191
Beryllium ..	Be	9	Oxygen ..	O	16
Bismuth ..	Bi	207·5	Palladium ..	Pd	106
Boron ..	B	11	Phosphorus ..	P	31
Bromine ..	Br	80	Platinum ..	Pt	195
Cadmium ..	Cd	112	Potassium ..	K	39
Cæsium ..	Cs	133	Rhodium ..	Rh	104
Calcium ..	Ca	40	Rubidium ..	Rb	85
Carbon ..	C	12	Ruthenium ..	Ru	103·5
Cerium ..	Ce	141	Samarium ..	Sm	150
Chlorine ..	Cl	35·5	Scandium ..	Sc	44
Chromium ..	Cr	52	Selenium (se-		
Cobalt ..	Co	59	lenion) ..	Se	79
Copper ..	Cu	63	Silicon ..	Si	28
Didymium ..	Di	145	Silver.. ..	Ag	108
Erbium ..	Er	166	Sodium ..	Na	23
Fluorine ..	F	19	Strontium ..	Sr	87
Gallium ..	Ga	70	Sulphur ..	S	32
Germanium ..	Ge	72	Tantalum ..	Ta	182
Gold ..	Au	197	Tellurium ..	Te	125
Hydrogen ..	H	1	Thallium ..	Tl	204
Indium ..	In	113	Thorium ..	Th	232
Iodine ..	I	127	Tin	Sn	118
Iridium ..	Ir	192·5	Titanium ..	Ti	48
Iron ..	Fe	56	Tungsten ..	W	184
Lanthanum ..	La	138·5	Uranium ..	U	240
Lead	Pb	207	Vanadium ..	V	51
Lithium ..	Li	7	Ytterbium ..	Yb	173
Magnesium ..	Mg	24	Yttrium ..	Y	90
Manganese ..	Mn	55	Zinc	Zn	65
Mercury ..	Hg	200	Zirconium ..	Zr	90
Molybdenum..	Mo	96			

Several molecules are represented by a large numeral placed in front of the formula; thus, $2\text{H}_2\text{SO}_4$ means 2 molecules of sulphuric acid, whilst portions of the molecules to be duplicated, etc., are usually bracketed, and a foot figure used; thus, $\text{Cu}(\text{NO}_3)_2$.

HOW FORMULÆ ARE FIXED.

Actual experiment enables us to find the percentage of the elements present in a particular compound. Thus, Expt. 129 informed us that carbon dioxide contained 27.27 per cent. of carbon and 72.73 per cent. of oxygen. Now, since the relative weights of the atoms of carbon and oxygen are 12 and 16, it follows that the **relative number of atoms** in carbon dioxide **will be found by dividing the percentage of each element by its atomic weight**; thus:

Relative number of

$$\begin{aligned} \text{C atoms} &= \frac{27.27}{12} = 2.27 \left\{ \begin{array}{l} \text{or, dividing} \\ \text{by the} \end{array} \right\} 1, \\ \text{O atoms} &= \frac{72.73}{16} = 4.54 \left\{ \begin{array}{l} \text{smaller} \\ (2.27) \end{array} \right\} \frac{4.54}{2.27}; \text{ i.e., } 2. \end{aligned}$$

Hence, the *simplest* formula (the **empirical formula**) contains carbon and oxygen in the proportion of 1 atom of the former to 2 atoms of the latter, and hence is written CO_2 .

Since the vapour density of carbon dioxide is 22 (Expt. 127), its molecular weight is $2 \times 22 = 44$.

The molecular formula is $(\text{CO}_2)_n$, which must have a weight, 44.

$$\text{C} = 12, \text{O}_2 = 2(16); \therefore \text{CO}_2 = 44.$$

Hence $(44)n = 44$; $\therefore n = 1$, and the **molecular formula** = CO_2 .

TO DETERMINE THE PERCENTAGE COMPOSITION FROM THE FORMULA.

Let us take, for example, CuO . Since the atomic weights of copper and oxygen are 63 and 16, $\text{CuO} = 63 + 16 = 79$.

So that 79 of copper oxide contain 63 of copper by weight;

\therefore 1 of copper oxide contains $\frac{63}{79}$ of copper by weight,

and 100 of copper oxide contain

$$\frac{63}{79} \times 100; \text{ i.e., } 79.75 \text{ per cent. of copper by weight.}$$

Similarly, copper oxide contains

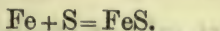
$$\frac{16}{79} \times 100; \text{ i.e., } 20.25 \text{ per cent. of oxygen by weight.}$$

Chemical Equations.

By means of these symbols and formulæ chemists are enabled *accurately* and *concisely* to *represent chemical actions* by means of *chemical equations* in their *quantitative* relations. To these equations the laws of chemical combination apply, and the two sides of the equation **must always balance**—i.e., every atom appearing on one side of the equation must occur on the other. *Before leaving an equation, it must always be verified to see that this is so.*

If we consider the reaction of Expt. 66, the equation representing the change is written: $\text{Mg} + \text{H}_2\text{O} = \text{MgO} + \text{H}_2$; we note at once that the equation does **not** tell us anything about “*how to bring the change about*” or “*about the nature of the action and products,*” etc.—i.e., a chemical equation **fails to state the conditions** of a chemical reaction.

Let us consider an equation in detail—*e.g.*,



We learn—

1. That iron and sulphur unite to form iron sulphide.

2. That 1 atom of iron weighing 56 unites with 1 atom of sulphur weighing 32 to form 1 molecule of iron sulphide weighing 88 (made up of 1 atom of sulphur, 16, and 1 atom of iron, 56).

Thus, from this equation we could calculate—

1. The weight of iron required to make 50 lbs. of iron sulphide.

For 88 lbs. of iron sulphide require 56 lbs. of iron;

\therefore 1 lb. of iron sulphide requires $\frac{56}{88}$ lbs. of iron,

and 50 lbs. of iron sulphide require

$$\frac{56}{88} \times 50; \text{ i.e., } 31.8 \text{ lbs. of iron.}$$

2. The weight of sulphur required to make 45 gms. of iron sulphide.

For 88 gms. of iron sulphide require 32 gms. of sulphur;

\therefore 1 gm. of iron sulphide requires $\frac{32}{88}$ gms. of sulphur;

\therefore 45 gms. of iron sulphide require

$$\frac{32}{88} \times 45; \text{ i.e., } 16.36 \text{ gms. of sulphur.}$$

3. The weight of iron sulphide from 3 tons of iron.

56 tons of iron produce 88 tons of iron sulphide;

1 ton of iron produces $\frac{88}{56}$ tons of iron sulphide;

3 tons of iron produce

$$\frac{88}{56} \times 3; \text{ i.e., } 4.71 \text{ tons of iron sulphide.}$$

4. The weight of iron sulphide from 65 ounces of sulphur.

32 ounces of sulphur produce 88 ounces of iron sulphide;

∴ 1 ounce of sulphur produces $\frac{88}{32}$ ounces of iron sulphide;

∴ 65 ounces of sulphur produce

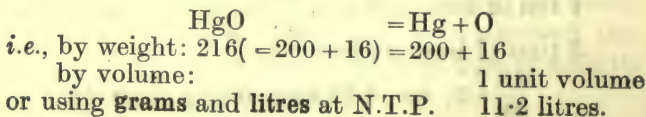
$$\frac{88}{32} \times 65; \text{ i.e., } 178.75 \text{ ounces of iron sulphide.}$$

In the above equation no gas is concerned. Now, it is found that if we use the **gram** as our unit of weight, and the **litre** as our unit of volume, that 1 gm. of hydrogen—i.e., the atomic weight of hydrogen in grams, or the so-called **gram atomic weight**) under 760 mm. pressure, and at 0° C. (the normal or standard temperature and pressure, often represented as N.T.P., or S.T.P.) occupies a volume of 11.2 litres, and this is true for the **gram atomic weight of all elementary gases**, and is termed the **gram atomic volume**.

Hence we may say that at N.T.P. the **gram atomic volume**—i.e., the volume of the atomic weight in grams—is 11.2 litres.

Since the molecule of hydrogen consists of 2 atoms, it has a gram molecular weight of 2, and hence a volume of 22.4 litres. So that the gram molecular weight of hydrogen at N.T.P. (and of all other gases) has the volume 22.4 litres; or at N.T.P. the **gram molecular volume**—i.e., the volume of the molecular weight in grams—is 22.4 litres.

Let us consider the equation—



We are now in a position to calculate as before—

- (1) Weight of mercury obtained from 50 lbs. of mercuric oxide.
- (2) Weight of oxygen obtained from 45 gms. of mercuric oxide.
- (3) Weight of mercuric oxide containing 3 tons of mercury.
- (4) Weight of mercuric oxide containing 65 ounces of oxygen.

And, in addition, we may determine—

5. The volume of oxygen at N.T.P. liberated from 50 gms. of mercuric oxide.

For 216 gms. of mercuric oxide liberate 11.2 litres of oxygen at N.T.P.;

∴ 1 gm. of mercuric oxide liberates $\frac{11.2}{216}$ litres of oxygen;

∴ 50 gms. of mercuric oxide liberate

$50 \times \frac{11.2}{216}$, or 2.59 litres of oxygen at N.T.P.

6. The weight of mercuric oxide liberating 15 litres of oxygen at N.T.P.

For 11.2 litres of oxygen at N.T.P. are obtained from 216 gms. of mercuric oxide;

∴ 1 litre of oxygen at N.T.P. is obtained from $\frac{216}{11.2}$ gms. of mercuric oxide;

and 15 litres of oxygen at N.T.P. are obtained from

$\frac{216}{11.2} \times 15$, or 289.3 gms. of mercuric oxide.

In this section is given a list of the reactions occurring in the experiments described in the preceding chapters. They are not intended to be memorized, but each equation should be read systematically, as

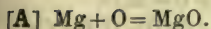
in the examples given below, and the experiment to which it refers should be recalled or revised. A valuable revision will thus be effected, while the greater number will become familiar to the student without much special effort. Moreover, a careful study will also reveal certain points of interest, some of which will be referred to later.

Chemical equations are commonly written in two ways—viz., atomic or molecular. In many cases chemical union takes place by *atoms uniting*, and thus an atomic equation represents the change at *the instant* the change occurs—*e.g.*, $\text{Mg} + \text{O} = \text{MgO}$ means that 1 atom of magnesium unites with 1 atom of oxygen to form 1 molecule of magnesium oxide. Rarely, however, do atoms exist free—*i.e.*, we begin with molecules, and when the change is completed, we end up with molecules—and so a molecular equation represents the initial and final stages of a chemical action, but some time before and after the change. The above action is thus expressed: $2\text{Mg} + \text{O}_2 = 2\text{MgO}$ —*i.e.*, 2 molecules of magnesium, containing 1 atom each, unite with 1 molecule of oxygen, consisting of 2 atoms, to produce 2 molecules of magnesium oxide.

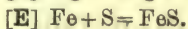
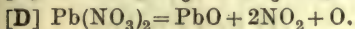
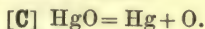
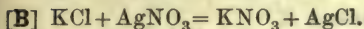
In some cases we do not know the number of atoms in a molecule, and a true molecular equation cannot be written—*e.g.*, $\text{C} + \text{O}_2 = \text{CO}_2$ is read thus: 1 atom of carbon unites with a molecule of oxygen to form a molecule of carbon dioxide. The equations below are written in their simplest form, and the student should rewrite some atomic equations in the molecular form, and *vice versa*.

TABLE OF EQUATIONS.

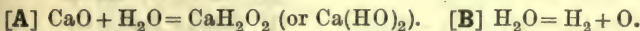
CHAPTER II.



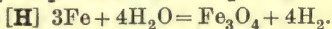
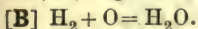
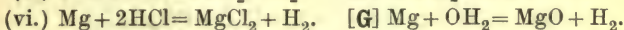
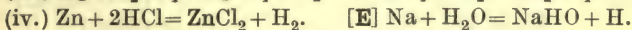
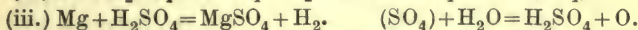
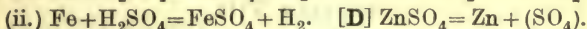
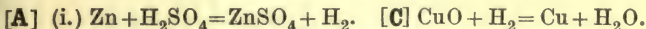
One atom of magnesium unites with one atom of oxygen to produce one molecule of magnesium oxide.



CHAPTER V.

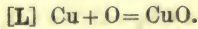
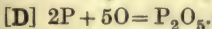
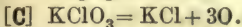
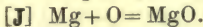
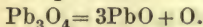
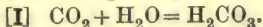
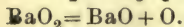
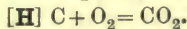
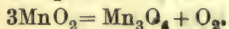
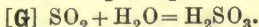
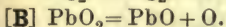
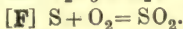
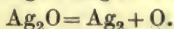
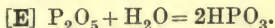
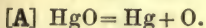


CHAPTER VI.

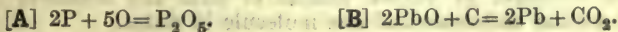


E.g., [A] (vi.) One atom of magnesium reacts with two molecules of hydrogen chloride to produce one molecule of magnesium chloride and one molecule of hydrogen.

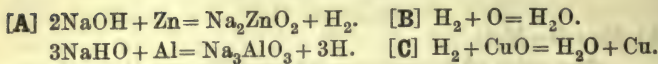
CHAPTER VII.



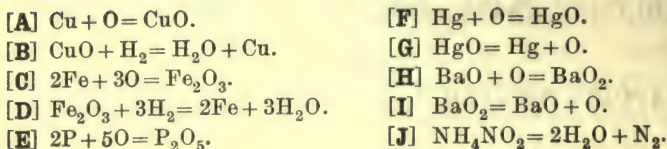
CHAPTER VIII.



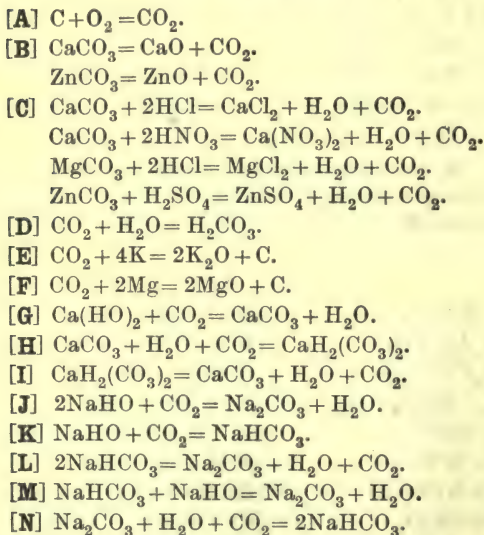
CHAPTER IX.



CHAPTER X.



CHAPTER XIII.



E.g., [J] two molecules of sodium hydroxide unite with one molecule of carbon dioxide to form one molecule of sodium carbonate and one molecule of water.

CHAPTER XIII.—Continued.

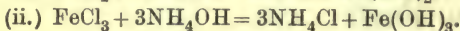
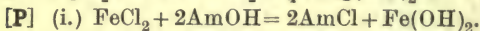
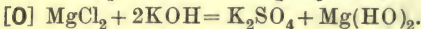
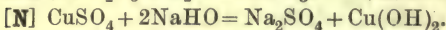
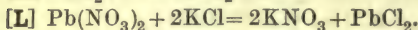
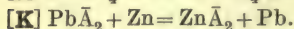
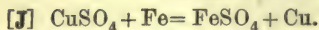
- [O] (i.) $2\text{KHO} + \text{CO}_2 = \text{K}_2\text{CO}_3 + \text{H}_2\text{O}$.
 (ii.) $\text{KHO} + \text{CO}_2 = \text{KHCO}_3$.
 (iii.) $2\text{KHCO}_3 = \text{K}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2$.
 (iv.) $\text{K}_2\text{CO}_3 + \text{H}_2\text{O} + \text{CO}_2 = 2\text{KHCO}_3$.
 [P] $\text{Na}_2\text{CO}_3 + \text{Ca}(\text{HO})_2 = \text{CaCO}_3 + 2\text{NaHO}$.
 [Q] $\text{K}_2\text{CO}_3 + \text{Ca}(\text{HO})_2 = \text{CaCO}_3 + 2\text{KHO}$.
 [R] $\text{CaCl}_2 + \text{K}_2\text{CO}_3 = \text{CaCO}_3 + 2\text{KCl}$.
 $\text{Ca}(\text{NO}_3)_2 + \text{K}_2\text{CO}_3 = \text{CaCO}_3 + 2\text{KNO}_3$.
 [S] $\text{CuSO}_4 + \text{Na}_2\text{CO}_3 = \text{CuCO}_3 + \text{Na}_2\text{SO}_4$.
 [T] $\text{CaH}_2(\text{CO}_3)_2 + \text{CaO} = 2\text{CaCO}_3 + \text{H}_2\text{O}$.

CHAPTER XIV.

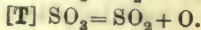
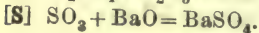
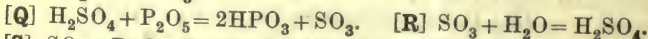
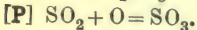
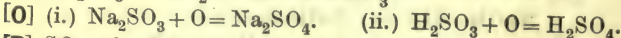
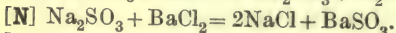
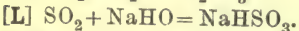
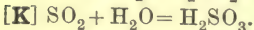
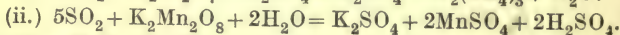
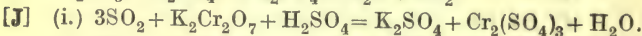
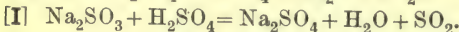
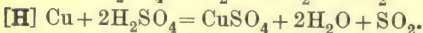
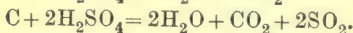
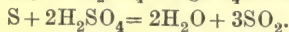
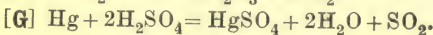
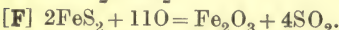
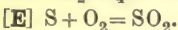
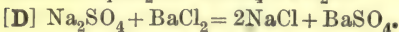
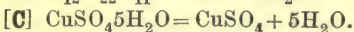
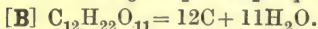
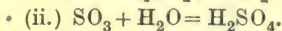
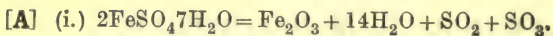
- [A] $\text{CO}_2 + \text{C} = 2\text{CO}$. [B] $\text{H}_2\text{O} + \text{C} = \text{CO} + \text{H}_2$.
 [C] $\text{NaHCO}_2 + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{H}_2\text{O} + \text{CO}$.
 $\text{Cu}(\text{HCO}_2)_2 + \text{H}_2\text{SO}_4 = \text{CuSO}_4 + 2\text{H}_2\text{O} + 2\text{CO}$.
 [D] $\text{H}_2\text{C}_2\text{O}_4 + \text{H}_2\text{SO}_4 = \text{H}_2\text{SO}_4 + \text{H}_2\text{O} + \text{CO}_2 + \text{CO}$.
 $\text{Na}_2\text{C}_2\text{O}_4 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + \text{H}_2\text{O} + \text{CO}_2 + \text{CO}$.
 [E] $\text{CO} + \text{O} = \text{CO}_2$. [F] $3\text{CO} + \text{Fe}_2\text{O}_3 = 2\text{Fe} + 3\text{CO}_2$.
 [G] $\text{CO} + \text{CuO} = \text{Cu} + \text{CO}_2$.

CHAPTER XVI.

- [A] $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$.
 [B] $3\text{Pb} + 8\text{HNO}_3 = 3\text{Pb}(\text{NO}_3)_2 + 4\text{H}_2\text{O} + 2\text{NO}$.
 [C] $\text{NaOH} + \text{HCl} = \text{NaCl} + \text{H}_2\text{O}$.
 [D] $\text{KHO} + \text{HNO}_3 = \text{KNO}_3 + \text{H}_2\text{O}$.
 [E] $\text{CuO} + \text{H}_2\text{SO}_4 = \text{CuSO}_4 + \text{H}_2\text{O}$.
 [F] $\text{MgO} + 2\text{HNO}_3 = \text{Mg}(\text{NO}_3)_2 + \text{H}_2\text{O}$.
 [G] $\text{CaCO}_3 + 2\text{HCl} = \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$.
 [H] $2\text{NaHO} + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$.
 $\text{NaOH} + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{H}_2\text{O}$.
 [I] (i.) $\text{H}_3\text{PO}_4 + \text{NaHO} = \text{NaH}_2\text{PO}_4 + \text{H}_2\text{O}$.
 (ii.) $\text{H}_3\text{PO}_4 + 2\text{NaHO} = \text{Na}_2\text{HPO}_4 + 2\text{H}_2\text{O}$.
 (iii.) $\text{H}_3\text{PO}_4 + 3\text{NaHO} = \text{Na}_3\text{PO}_4 + 3\text{H}_2\text{O}$.

CHAPTER XVI.—*Continued.*

CHAPTER XVII.



CHAPTER XVII.—Continued.

- [U] (i.) $\text{SO}_2 + 2\text{HNO}_3 = \text{H}_2\text{SO}_4 + 2\text{NO}_2$.
 (ii.) $3\text{SO}_2 + 2\text{H}_2\text{O} + 2\text{HNO}_3 = 3\text{H}_2\text{SO}_4 + 2\text{NO}$.
 (iii.) $\text{NO} + \text{O} = \text{NO}_2$.
 (iv.) $\text{NO}_2 + \text{SO}_2 = \text{NO} + \text{SO}_3$. (v.) $\text{SO}_3 + \text{H}_2\text{O} = \text{H}_2\text{SO}_4$.
 [V] (i.) $2\text{SO}_2 + \text{H}_2\text{O} + 3\text{NO}_2 = 2\text{H}(\text{NO})\text{SO}_4 + \text{NO}$.
 (ii.) $2\text{H}(\text{NO})\text{SO}_4 + \text{H}_2\text{O} = 2\text{H}_2\text{SO}_4 + \text{NO} + \text{NO}_2$.

CHAPTERS XVIII. AND XIX.

- [A] $3\text{FeS}_2 = \text{Fe}_3\text{S}_4 + \text{S}_2$. [B] $\text{S} + \text{O}_2 = \text{SO}_2$.
 [C] $\text{S} + \text{H}_2 = \text{SH}_2$. [D] $2\text{Cu} + \text{S} = \text{Cu}_2\text{S}$.
 [E] $\text{Fe} + \text{S} = \text{FeS}$. [F] $\text{C} + \text{S}_2 = \text{CS}_2$.
 [G] $\text{Na}_2\text{S}_2\text{O}_3 + 2\text{HCl} = 2\text{NaCl} + \text{H}_2\text{O} + \text{SO}_2 + \text{S}$.
 [H] (i.) $\text{FeS} + 2\text{HCl} = \text{FeCl}_2 + \text{SH}_2$.
 (ii.) $\text{FeS} + \text{H}_2\text{SO}_4 = \text{FeSO}_4 + \text{H}_2\text{S}$.
 (iii.) $\text{CuS} + 2\text{HCl} = \text{CuCl}_2 + \text{H}_2\text{S}$.
 [J] $\text{H}_2\text{S} + \text{O} = \text{H}_2\text{O} + \text{S}$.
 [K] $\text{H}_2\text{S} + 3\text{O} = \text{H}_2\text{O} + \text{SO}_2$.
 [L] $2\text{H}_2\text{S} + \text{SO}_2 = 2\text{H}_2\text{O} + 3\text{S}$.
 [M] (i.) $\text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{S} + 4\text{H}_2\text{SO}_4 = \text{K}_2\text{SO}_4 + \text{Cr}_2(\text{SO}_4)_3 + 4\text{H}_2\text{O} + \text{H}_2\text{SO}_3$.
 (ii.) $4\text{K}_2\text{Cr}_2\text{O}_7 + 3\text{H}_2\text{S} + 13\text{H}_2\text{SO}_4 = 4\text{K}_2\text{SO}_4 + 4\text{Cr}_2(\text{SO}_4)_3 + 16\text{H}_2\text{O}$.
 [N] $4\text{K}_2\text{Mn}_2\text{O}_8 + 5\text{H}_2\text{S} + 7\text{H}_2\text{SO}_4 = 4\text{K}_2\text{SO}_4 + 8\text{MnSO}_4 + 12\text{H}_2\text{O}$.
 [O] $\text{H}_2\text{SO}_4 + \text{H}_2\text{S} = 2\text{H}_2\text{O} + \text{SO}_2 + \text{S}$.
 [P] $\text{K} + \text{H}_2\text{S} = \text{KHS} + \text{H}$. $\text{Sn} + \text{H}_2\text{S} = \text{SnS} + \text{H}_2$.
 $\text{Ag}_2 + \text{H}_2\text{S} = \text{Ag}_2\text{S} + \text{H}_2$.
 [Q] $\text{Pb}(\text{NO}_3)_2 + \text{H}_2\text{S} = 2\text{HNO}_3 + \text{PbS}$.
 $\text{Hg}(\text{NO}_3)_2 + \text{SH}_2 = 2\text{HNO}_3 + \text{HgS}$.
 $\text{CuSO}_4 + \text{H}_2\text{S} = \text{H}_2\text{SO}_4 + \text{CuS}$.
 $\text{CdCl}_2 + \text{SH}_2 = 2\text{HCl} + \text{CdS}$.
 $2\text{FeCl}_3 + 3\text{H}_2\text{S} = 2\text{FeS} + \text{S} + 6\text{HCl}$.
 [R] $\text{Sb}_2\text{S}_3 + 6\text{HCl} = 2\text{SbCl}_3 + 3\text{SH}_2$. [S] $\text{Fe} + \text{H}_2\text{S} = \text{FeS} + \text{H}_2$.

CHAPTER XX.

- [A] $\text{KNO}_3 + \text{H}_2\text{SO}_4 = \text{KHSO}_4 + \text{HNO}_3$.
- [B] $\text{Cu}(\text{NO}_3)_2 + \text{H}_2\text{SO}_4 = \text{CuSO}_4 + 2\text{HNO}_3$.
- [C] $2\text{NaNO}_3 + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{HNO}_3$.
- [D] (i.) $4\text{Na}_2\text{SO}_3 + 2\text{HNO}_3 = 4\text{Na}_2\text{SO}_4 + \text{H}_2\text{O} + 2\text{NO}$.
 (ii.) $3\text{S} + 8\text{HNO}_3 = 3\text{H}_2\text{SO}_4 + 4\text{H}_2\text{O} + 8\text{NO}$.
- [E] $10\text{FeSO}_4 + 2\text{HNO}_3 + 3\text{H}_2\text{SO}_4 = 3\text{Fe}_2(\text{SO}_4)_3 + 4\text{H}_2\text{O} + 2(2\text{FeSO}_4 \cdot \text{NO})$.
- [F] $\text{Mg} + 2\text{HNO}_3 = \text{Mg}(\text{NO}_3)_2 + \text{H}_2$.
- [G] $2\text{HNO}_3 = \text{H}_2\text{O} + 2\text{NO} + 3\text{O}$.
- [H] $\text{Cu}(\text{NO}_3)_2 = \text{CuO} + 2\text{NO}_2 + \text{O}$.
- [I] $\text{KNO}_3 = \text{KNO}_2 + \text{O}$.
- [J] (i.) $3\text{Cu} + 8\text{HNO}_3 = 3\text{Cu}(\text{NO}_3)_2 + 4\text{H}_2\text{O} + 2\text{NO}$.
 (ii.) $4\text{NO} + 3\text{Fe} = \text{Fe}_3\text{O}_4 + 2\text{N}_2$.
 (iii.) $\text{NO} + \text{Cu} = \text{CuO} + \text{N}$.
- [K] $\text{Pb}(\text{NO}_3)_2 = \text{PbO} + 2\text{NO}_2 + \text{O}$.
- [L] $6\text{FeSO}_4 + 3\text{H}_2\text{SO}_4 + 2\text{HNO}_3 = 3\text{Fe}_2(\text{SO}_4)_3 + 4\text{H}_2\text{O} + 2\text{NO}$.
- [M] $\text{NO} + \text{O} = \text{NO}_2$.
- [N] $2\text{P} + 5\text{NO} = \text{P}_2\text{O}_5 + 5\text{N}$.
- [O] $2\text{FeSO}_4 + \text{NO} = 2\text{FeSO}_4 \cdot \text{NO}$.
- [P] $\text{N}_2\text{O}_4 + \text{H}_2\text{O} = \text{HNO}_2 + \text{HNO}_3$.
- [Q] $3\text{NO}_2 + \text{H}_2\text{O} = 2\text{HNO}_3 + \text{NO}$.
- [R] $\text{Hg} + \text{NO}_2 = \text{HgO} + \text{NO}$.
- [S] (i.) $\text{P}_4 + 5\text{NO}_2 = 2\text{P}_2\text{O}_5 + 5\text{N}$. (ii.) $\text{S} + \text{NO}_2 = \text{SO}_2 + \text{N}$.
- [T] (i.) $3\text{Fe} + \text{N}_2\text{O}_4 = \text{Fe}_3\text{O}_4 + \text{N}_2$. (ii.) $2\text{Cu} + \text{NO}_2 = 2\text{CuO} + \text{N}$.
- [U] $\text{NH}_4\text{OH} + \text{HNO}_3 = \text{NH}_4\text{NO}_3 + \text{H}_2\text{O}$.
- [V] $\text{NH}_4\text{NO}_3 + \text{NaOH} = \text{NH}_3 + \text{H}_2\text{O} + \text{NaNO}_3$.
- [W] (i.) $\text{NH}_4\text{NO}_3 = 2\text{H}_2\text{O} + \text{N}_2\text{O}$.
 (ii.) $\text{NH}_4\text{NO}_3 = 2\text{H}_2\text{O} + \text{N} + \text{NO}$.
- [X] $5\text{N}_2\text{O} + 2\text{P} = \text{P}_2\text{O}_5 + 5\text{N}_2$.
- [Y] (i.) $4\text{N}_2\text{O} + 3\text{Fe} = \text{Fe}_3\text{O}_4 + 4\text{N}_2$. (ii.) $\text{N}_2\text{O} + \text{Cu} = \text{CuO} + \text{N}_2$.

CHAPTER XXI.

- [A] (i.) $2\text{NH}_4\text{Cl} + \text{Ca}(\text{OH})_2 = \text{CaCl}_2 + 2\text{H}_2\text{O} + 2\text{NH}_3$.
 (ii.) $2\text{NH}_4\text{Cl} + \text{CaO} = \text{CaCl}_2 + \text{H}_2\text{O} + 2\text{NH}_3$.
 (iii.) $\text{NH}_4\text{NO}_3 + \text{KHO} = \text{KNO}_3 + \text{H}_2\text{O} + \text{NH}_3$.

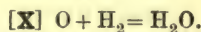
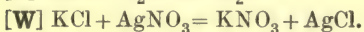
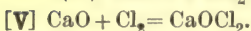
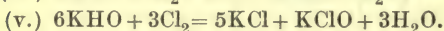
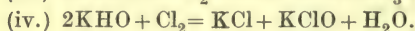
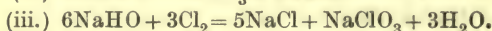
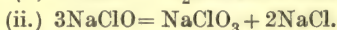
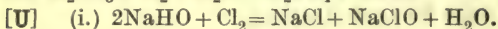
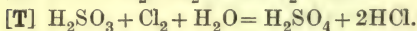
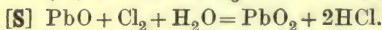
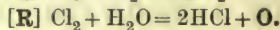
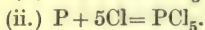
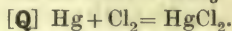
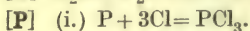
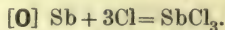
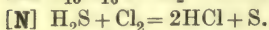
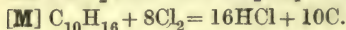
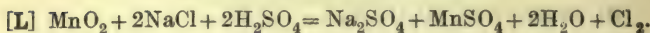
CHAPTER XXI.—*Continued.*

- [B] $2\text{NH}_3 + 3\text{O} = 3\text{H}_2\text{O} + \text{N}_2$.
- [C] $\text{CaCl}_2 + 4\text{NH}_3 = \text{CaCl}_2 \cdot 4\text{NH}_3$.
- [D] $\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}$.
- [E] $\text{NH}_4\text{Cl} + \text{NaHO} = \text{NaCl} + \text{H}_2\text{O} + \text{NH}_3$.
- [F] (i.) $\text{CuSO}_4 + 2\text{NH}_4\text{OH} = \text{Cu}(\text{OH})_2 + (\text{NH}_4)_2\text{SO}_4$.
 (ii.) $\text{CuSO}_4 + 2\text{NaOH} = \text{Cu}(\text{OH})_2 + \text{Na}_2\text{SO}_4$.
- [G] (i.) $\text{NH}_4\text{OH} + \text{HCl} = \text{NH}_4\text{Cl} + \text{H}_2\text{O}$.
 (ii.) $\text{KOH} + \text{HCl} = \text{KCl} + \text{H}_2\text{O}$.
 (iii.) $\text{NH}_4\text{OH} + \text{HNO}_3 = \text{NH}_4\text{NO}_3 + \text{H}_2\text{O}$.
 (iv.) $\text{NaOH} + \text{HNO}_3 = \text{NaNO}_3 + \text{H}_2\text{O}$.
 (v.) $2\text{NH}_4\text{OH} + \text{H}_2\text{SO}_4 = (\text{NH}_4)_2\text{SO}_4 + 2\text{H}_2\text{O}$.
 (vi.) $2\text{KOH} + \text{H}_2\text{SO}_4 = \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$.
 (vii.) $\text{NH}_4\text{OH} + \text{H}_2\text{SO}_4 = (\text{NH}_4)\text{HSO}_4 + \text{H}_2\text{O}$.
 (viii.) $\text{NaOH} + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{H}_2\text{O}$.
- [H] $\text{NH}_3 + \text{H}_2\text{O} = \text{NH}_4\cdot\text{OH}$.
- [I] $2\text{NH}_3 + 3\text{CuO} = 3\text{H}_2\text{O} + 3\text{Cu} + \text{N}_2$.
- [J] $\text{NH}_3 + 3\text{Cl} = 3\text{HCl} + \text{N}$.

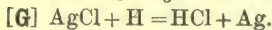
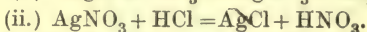
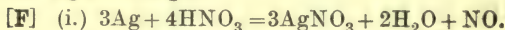
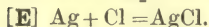
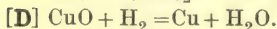
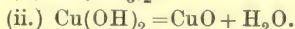
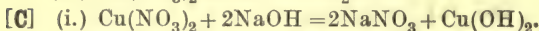
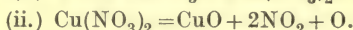
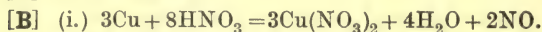
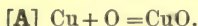
CHAPTER XXII.

- [A] (i.) $\text{NaCl} + \text{H}_2\text{SO}_4 = \text{NaHSO}_4 + \text{HCl}$.
 (ii.) $\text{CuCl}_2 + \text{H}_2\text{SO}_4 = \text{CuSO}_4 + 2\text{HCl}$.
- [B] $2\text{NaCl} + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{HCl}$.
- [C] $\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}$.
- [D] (i.) $\text{Zn} + 2\text{HCl} = \text{ZnCl}_2 + \text{H}_2$.
 (ii.) $\text{Mg} + 2\text{HCl} = \text{MgCl}_2 + \text{H}_2$.
- [E] (i.) $\text{Al} + 3\text{HCl} = \text{AlCl}_3 + 3\text{H}$.
 (ii.) $\text{Fe} + 2\text{HCl} = \text{FeCl}_2 + 2\text{H}$.
 (iii.) $\text{Na} + \text{HCl} = \text{NaCl} + \text{H}$.
- [F] $4\text{HCl} + \text{MnO}_2 = 2\text{H}_2\text{O} + \text{MnCl}_2 + \text{Cl}_2$.
- [G] $\text{H} + \text{Cl} = \text{HCl}$. [H] $\text{HCl} = \text{H} + \text{Cl}$.
- [I] $\text{AgCl} + \text{H} = \text{HCl} + \text{Ag}$.
- [J] $\text{HNO}_3 + 3\text{HCl} = 2\text{H}_2\text{O} + \text{NOCl} + \text{Cl}_2$.
- [K] (i.) $\text{MnO}_2 + 4\text{HCl} = 2\text{H}_2\text{O} + \text{MnCl}_2 + \text{Cl}_2$.
 (ii.) $\text{PbO}_2 + 4\text{HCl}_2 = 2\text{H}_2\text{O} + \text{PbCl}_2 + \text{Cl}_2$.

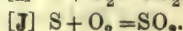
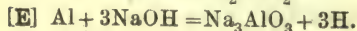
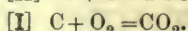
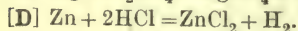
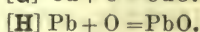
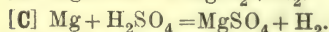
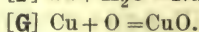
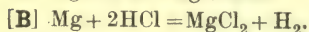
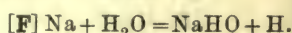
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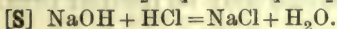
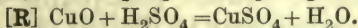
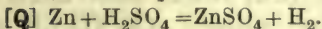
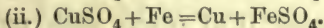
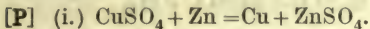
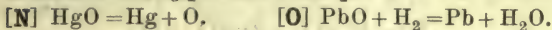
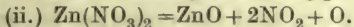
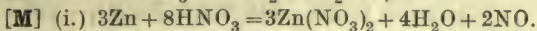
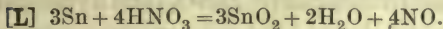


CHAPTER XXIII.



CHAPTER XXIV.



CHAPTER XXIV.—*Continued.*

CHAPTER XXV.



A careful study of the above equations will inform us of many interesting points. It must, however, be remembered that an **equation follows the experiment** which it represents; the experiment is not performed after an equation is evolved. Equations are therefore *not to be remembered* as a collection of symbols and formulæ, but *are to be built up* from the results of the experiment. In reading the above equations, therefore, use the systematic names, as in the examples given.

EQUATION BUILDING.

Before being able to build up an equation, we must know (1) the bodies reacting and their formulæ; (2) the bodies produced and their formulæ.

EXAMPLE 1.—Sulphur (S) in Oxygen (O) produces sulphur dioxide (SO_2).

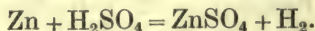
∴ Equation is $\text{S} + x\text{O} = \text{SO}_2.$

Obviously x is 2, and the correct equation becomes $\text{S} + \text{O}_2 = \text{SO}_2.$

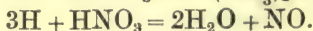
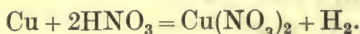
EXAMPLE 2.—Zinc (Zn) and sulphuric acid (H_2SO_4) react to produce zinc sulphate (ZnSO_4) and hydrogen (H).

∴ Equation is $\text{Zn} + x\text{H}_2\text{SO}_4 = \text{ZnSO}_4 + y\text{H}$.

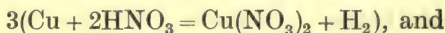
Obviously $x = 1$, $y = 2$, and the equation becomes



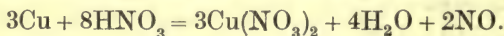
EXAMPLE 3.—Copper (Cu) reacts with nitric acid (HNO_3), forming copper nitrate ($\text{Cu}(\text{NO}_3)_2$), nitric oxide (NO). This equation may be derived thus:



Now, since hydrogen does not escape, these two actions must occur in such proportions that the hydrogen liberated in 1 is used up in 2—*i.e.*,



Simple addition gives us the final equation—



Several points which can be gathered from a study of the above equations are of great assistance in the making of equations.

A. 1. The atoms of the metals sodium, potassium, silver, ammonium, always replace *one* atom of hydrogen.

2. The atoms of most other metals, copper, lead, magnesium, zinc, iron, etc., replace *two* atoms of hydrogen.

3. The atoms of aluminium, and occasionally iron (ferric) replace *three* atoms of hydrogen.

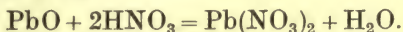
B. 1. The atoms of the non-metal chlorine combine with one atom of hydrogen.

2. The atoms of the non-metals, oxygen, sulphur, etc., combine with two atoms of hydrogen.

3. The atoms of nitrogen and phosphorus with three atoms.

If these are remembered, many formulæ need not be committed to memory, for in reactions between these elements, 1 atom of the metals in *A* 1 unite with 1 atom of the non-metals in *B* 1; 1 atom of the metals in *A* 2 unite with 1 atom of the non-metals in *B* 2, and so on. Further, 2 atoms of the metals in *A* 1 will unite with 1 atom of the non-metals in *B* 2, or 2 atoms of the non-metals in *B* 1 will unite with 1 atom of the metals in *A* 2, etc. Thus a knowledge of a few of the simpler formulæ will suffice.

EXAMPLE 1.—Write down the equation for the action of lead oxide on nitric acid. One atom of lead replaces 2 atoms of hydrogen, and 1 atom of oxygen unites with 2 atoms of hydrogen; therefore the formula of lead oxide is PbO . The formula of nitric acid is HNO_3 ; and since 1 atom of lead replaces 2 atoms of hydrogen, the equation must be—



Thus, for this example the *only* formula to be *remembered* is HNO_3 .

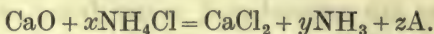
EXAMPLE II.—Write down the equation expressing the reaction between calcium oxide and ammonium chloride.

Here calcium and oxygen belong to Class II. The formula of calcium oxide is therefore CaO .

Ammonium and chlorine belong to Class I. The formula of ammonium chloride is therefore NH_4Cl .

Since calcium is of Class II. and chlorine of Class I., the formula of calcium chloride is CaCl_2 , and we *know* that ammonia (NH_3) is liberated; hence the equation

can be written at once, keeping the number of atoms of calcium the same on both sides:



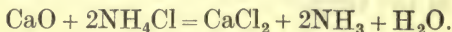
x is obviously 2 (Why ?); hence y is 2. (What is $z\text{A}$?)

Writing the equation thus:



we find 1 atom of oxygen and 2 atoms of hydrogen are missing from the right-hand side. Hence A is water and z is 1.

Thus, the completed equation is—



With the aid of knowledge thus gained from these simple changes, even complicated equations may be built up. The student is advised to try to do this for all reactions which are not quite easy, and endeavour to discover the *keynote*, or true course, of the reaction. The labour will be well spent, and the reward more than adequate.

CHAPTER XXVII

BOYLE'S LAW—CHARLES'S LAW—CHEMICAL CALCULATIONS

IN all the experiments in the preceding sections we were careful to read the volumes of gases under the same temperature and pressure throughout the experiment.

Rarely, however, during experiments is the barometer at 760 mm. or the temperature at 0°C. , so that, if we wish to determine the volume of a gas obtainable as in Chapter XXV. under any given set of conditions, we must investigate the behaviour of gases with respect to pressure and temperature.

EXPT. 290.—To show that Gases exert Pressure.—

1. Tie a piece of thin paper over the air-pump receiver shown (Fig. 92), and withdraw the air from below. Very soon the paper gives way with a sharp report. (Why does it collapse?)

2. Partially fill a toy india-rubber balloon, and firmly tie the neck with thread. (Why does it remain expanded?) Place this under the receiver of an air-pump, and exhaust as before. (Why does it expand and ultimately burst?)

3. Completely fill a tumbler with water and cover

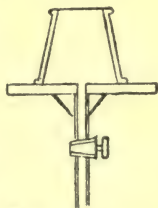


FIG. 92.

it with a piece of paper. Holding the paper on firmly, invert it and remove the hand. Why does the paper remain on and no water escape ?

4. Fill a narrow glass tube, about 33 inches long, with mercury ; close the end with the thumb, and invert it in a trough of mercury. Observe that the



FIG. 93.

mercury falls. (What is in the space above the mercury ?) Gradually slope the tube as at *B*, and notice that the mercury is always at the same vertical height. The tube becomes quite full of mercury at *C*—i.e., the space above the mercury is empty ; i.e., is a **vacuum**. (What, then, keeps the mercury always at the same level ?) It is obviously the *air* pressing on the mercury in the trough that supports the mercury column. Such an

apparatus is known as a **barometer**, or pressure measurer. (What does its movement from day to day really mean ? Why do gas jars remain full of water in the trough when we collect gases ?)

When the air is in motion, we feel it pressing on and about us, and we speak of it as “wind.” Other gases behave in exactly the same way.

EXPT. 291.—To study the Behaviour of Gas with Reference to Pressure.—Fit up the apparatus shown (Fig. 94). It consists of a graduated gas-tube (*A*) attached to a reservoir-tube (*B*) by a long length of india-rubber tubing. Open the tap, and holding the tubes

together, pour mercury into *B* until it stands half-up in *A*. Close the tap and read the volume. The mercury is at the same level in both tubes. (What is the pressure on *A*?) Now fix *A* in position, and raise *B* 1 foot, 2 feet, etc., above the mercury-level in *A*, and read (1) the volume of the gas, (2) the height of the mercury in *B* above the mercury in *A*. Lower *B* below *A*, and again take these readings. Repeat this experiment, filling *A* with coal-gas, hydrogen, carbon dioxide, etc. Do we not learn—

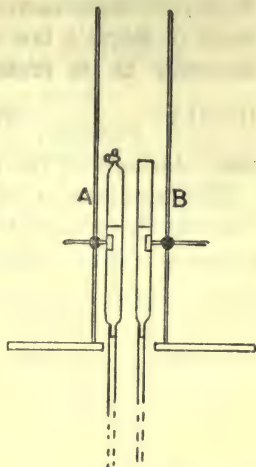


FIG. 94.

(1) That all gases behave alike with regard to pressure?

(2) That the volume of gas at constant temperature depends upon the pressure? It is greater when the pressure is low, and less when the pressure is high. Tabulate your results thus:

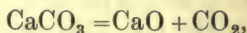
BAROMETER = 76 CMS.

	Volume (V).	Height of B.	Pressure (P).	P. × V.
A.	12.0 c.cms.	0.0 cms.	76.0 + 0.0 = 76.0	912
	9.0 "	25.0 "	76.0 + 25.0 = 101.0	909
	7.5 "	46.0 "	76.0 + 46.0 = 122.0	915
	6.0 "	75.0 "	76.0 + 75.0 = 151.0	906
B.	15.0 "	15.0 "	76.0 - 15.0 = 61.0	915
	17.5 "	24.0 "	76.0 - 24.0 = 52.0	910
	20.0 "	30.5 "	76.0 - 30.5 = 45.5	910
Average				911

Thus, at the same temperature the product **PV** is a constant, or $P_1V_1 = P_2V_2$. This may be expressed, as Boyle, its discoverer, expressed it in 1650, in his statement of **Boyle's law**—viz., **The volume of a gas varies inversely as its pressure**—i.e., the volume is proportional to $\frac{1}{\text{pressure}}$, provided the temperature remains the same throughout the experiment.

The application of Boyle's law to calculations involving the volumes of gases is best illustrated by an example. For instance, suppose that it is required to find the volume of gas evolved by 70 gms. of chalk at 0° C. and 800 mm. pressure.

From the equation



i.e., by weight

$$100 \quad 56 + 44$$

by volume (grams, litres) at N.T.P. 22.4

we find that 100 gms. of chalk liberate 22.4 litres at N.T.P.;

hence 1 gm. of chalk liberates $\frac{22.4}{100}$ litres at N.T.P.,

and 70 gms. of chalk liberate

$$\frac{22.4}{100} \times 70, \text{ or } 15.68 \text{ litres, at N.T.P.}$$

Since the pressure in this experiment is greater than the normal, the volume of the gas will be less than that above. Hence, at 0° C. and 800 mm. the volume is

$15.68 \times \frac{760}{800}$ litres—i.e., **14.90 litres of carbon dioxide** will be obtained.

Now, supposing we wish to obtain the weight of chalk that will yield 12 litres of carbon dioxide at 0° C. and 900 mm. pressure. Since the equation only informs us of the volume of the gas at 0° C. and

760 mm., we first of all find what 12 litres at 900 mm. pressure would be at 760 mm. Since the **pressure is to be reduced** from 900 mm. to 760 mm., the **volume will increase**, and becomes $\frac{12 \times 900}{760}$, or 14.21 litres at S.T.P.

Now, from the equation we find that

22.4 litres of carbon dioxide are obtained from 100 gms. of chalk;

\therefore 1 litre of carbon dioxide is obtained from $\frac{100}{22.4}$ gms. of chalk;

\therefore 14.21 litres of carbon dioxide are obtained from

$$\frac{100}{22.4} \times 14.21; \text{ i.e., } 63.44 \text{ gms. of chalk.}$$

Lastly, suppose we desire to know the volume of carbon dioxide given off by 40 gms. of chalk at 0° C. and 600 mm. pressure.

The equation states that

100 gms. of chalk give off 22.4 litres of carbon dioxide at N.T.P.;

\therefore 1 gm. of chalk gives off $\frac{22.4}{100}$ litres of carbon dioxide at N.T.P.,

and 40 gms. of chalk give off

$$\frac{22.4}{100} \times 40, \text{ or } 8.96 \text{ litres of carbon dioxide, at N.T.P.}$$

Now, as the experimental pressure, 600 mm., is less than normal pressure, 760 mm., the **volume will be greater**. Hence, at 0° C. and 600 mm. pressure the volume becomes $8.96 \times \frac{760}{600} = 11.35$ litres.

Let us next investigate the variation of volume with temperature.

EXPT. 292.—To study the Behaviour of Gases with respect to Temperature.—As we learned in Expt. 7, all

gases expand when heated, and contract when cooled, and, as far as we could judge there, they expanded regularly and equally. Let us make a closer examination. Take a flask (about 250 c.cms.) fitted with a cork, tube, and plug, as shown (Fig. 49), and fill it with cold water. Pour this out into a graduated cylinder, and note its volume. Thoroughly dry the flask, remove the plug, and immerse it in a bath of boiling water for about ten minutes. The flask is now full of air at 100°C . Replace the plug firmly, remove the flask from the bath, and open it under ice-cold water, completely immersing the flask. When no more water enters, withdraw it until the water inside and outside are level (Why ?), plug the opening, and determine the volume of the water that has entered.

The volume of flask $= 250$ c.cms.;
the volume of water that entered $= 67$ c.cms.

Hence at 100°C . the air in the flask had a volume of 250 c.cms.,

and at 0°C . the same air in the flask had a volume of $250 - 67 = 183$ c.cms.

So that 183 c.cms. of air at 0°C . must have increased in volume 67 c.cms. when heated through 100°C .;

\therefore 1 c.cm. of air at 0°C . must have increased in volume

$\frac{67}{183}$ c.cms. when heated through 100°C .,

and 1 c.cm. of air at 0°C . must have increased in volume

$\frac{67}{183 \times 100}$ c.cms. when heated through 1°C .;

i.e., $\frac{1}{276}$.

Repeat this experiment with air, warming it to 60°C ., 40°C ., 20°C . and again with other gases—

hydrogen, coal-gas, nitrogen, oxygen, etc. In all cases, if accurately carried out, we find (1) gases expand regularly; (2) all gases expand equally when heated through the same range of temperature.

Charles, who investigated this carefully, found that at constant pressure all gases expand $\frac{1}{273}$ of their volume at 0°C . for every degree rise of temperature. This is known as **Charles's law**.

ABSOLUTE TEMPERATURE.—A gas having a volume of 273 c.cms. at 0°C . would, if cooled to -1°C ., have a volume of 272 c.cms.; at -2°C ., a volume of 271 c.cms., and so on. If the law held, then, at -273°C ., its volume would be 0. Of course, long before this temperature is reached the gases become liquid, and the law breaks down. Nevertheless, this temperature, -273°C ., is taken as the zero (the **absolute zero**) on a new scale of temperatures, termed **absolute temperatures**.

COMPARISON OF CENTIGRADE AND ABSOLUTE TEMPERATURES.—It is obvious from a study of the accom-



FIG. 95.

panying diagram that the absolute temperature = centigrade temperature + 273.

$$\text{Thus, } t^{\circ}\text{C.} = (273 + t)^{\circ}\text{A.}$$

We may now express Charles's law very simply with respect to this scale of absolute temperature, thus:

Since by Charles's law

1 vol. of gas at 0° C. becomes $1 + \left(\frac{1}{273}\right)t$ vols. at t° C.;

i.e., V_0 of gas at 0° C. becomes $V_0 \left(1 + \frac{t}{273}\right)$ vols. at t° C.

i.e., volume at t° C. $V_t = V_0 \left(\frac{273+t}{273}\right)$,

$$= V_0 \left(\frac{273+t}{273+0}\right),$$

$$= V_0 \left(\frac{A_t}{A_0}\right);$$

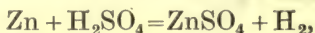
$$\text{or } \frac{V_t}{V_0} = \frac{A_t}{A_0};$$

i.e., the volume of a gas is proportional to the absolute temperature.

We are now able to apply this law to the solution of new problems.

1. Required the weight of zinc that, when dissolved in dilute sulphuric acid, will liberate 35 litres of hydrogen at 15° C. and 760 mm.

The equation expressing the change is—



i.e., by weight,

$$65 + 98 = 161 + 2,$$

by volume at N.T.P.

22.4 litres.

Now, since the equation only gives us volumes at 0° C., and we want the gas at 15° C., we must calculate what the required volume would be at 0° C.

$$0^{\circ} \text{ C.} = 273^{\circ} \text{ A.} \quad 15^{\circ} \text{ C.} = 273 + 15 = 288^{\circ} \text{ A.}$$

Since the temperature to be considered— 273° A —is lower than the required temperature—*i.e.*, 288° A —the volume will be less—viz.,

$$35 \times \frac{273}{288}, \text{ or } 33.2 \text{ litres at } 0^{\circ} \text{ C.}$$

Now, the equation states

i.e., 22.4 litres of hydrogen at 0° C. and 760 mm. are obtained from 65 gms. of zinc;

\therefore 1 litre of hydrogen at 0° C. and 760 mm. is obtained from $\frac{65}{22.4}$ gms. of zinc,

and 33.2 litres of hydrogen at 0° C. and 760 mm. are obtained from $\frac{65}{22.4} \times 33.2$, or 96.34 gms. of zinc.

2. What volume of hydrogen will be obtained from 50 gms. of zinc at 20° C. and 760 mm. pressure?

$$0^{\circ} \text{ C.} = 273^{\circ} \text{ A.} \quad 20^{\circ} \text{ C.} = 273 + 20 = 293^{\circ} \text{ A.}$$

Since the gas is to be measured at a higher temperature, the volume will be greater.

Our equation states that 65 gms. of zinc give off 22.4 litres of hydrogen at 0° C. and 760 mm.; and hence the equation becomes

65 gms. of zinc give off $\frac{22.4 \times 293}{273}$ litres of hydrogen at

20° C. and 760 mm.; *i.e.*, 24.04 litres of hydrogen;

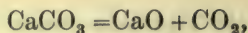
hence 1 gm. of zinc gives off $\frac{24.04}{65}$ litre of hydrogen at 20° C. and 760 mm.,

and 50 gms. of zinc give off $\frac{24.04}{65} \times 50$, or 18.5 litres of hydrogen at 20° C. and 760 mm.

When both the pressure and the temperature vary, the solution is carried out in the same way:

1. What volume of carbon dioxide at 20° C. and 850 mm. is obtained from 60 gms. of chalk when heated?

The equation,



i.e., by weight

$$100 = 56 + 44,$$

by volume at N.T.P.

22.4 litres,

states that 100 gms. of chalk give off 22.4 litres at N.T.P.

Since the **pressure is greater** than this, the **volume will be less**—viz.,

$$22.4 \times \frac{760}{850}, \text{ or } 20.03 \text{ litres at } 0^\circ \text{ C. and } 850 \text{ mm.}$$

$$0^\circ \text{ C. is } 273^\circ \text{ A. } 20^\circ \text{ C. } = 273 + 20 = 293^\circ \text{ A.}$$

Since the **temperature is greater** than normal temperature, the **volume will be greater**; viz.,

$$\frac{20.03 \times 293}{273}; \text{ i.e., } 21.5 \text{ litres at } 20^\circ \text{ C. and } 850 \text{ mm.}$$

So that 100 gms. of chalk liberate 21.5 at $20^\circ \text{ C. and } 850 \text{ mm.}$;

\therefore 1 gm. of chalk liberates $\frac{21.5}{100}$ at $20^\circ \text{ C. and } 850 \text{ mm.}$,
and 60 gms. of chalk liberate

$$\frac{21.5}{100} \times 60, \text{ or } 12.9 \text{ litres at } 20^\circ \text{ C. and } 850 \text{ mm.}$$

2. What weight of zinc is required to generate 27 litres of hydrogen at $15^\circ \text{ C. and } 740 \text{ mm.}$?

Before we can use our equation, we wish to know the volume of this gas at $0^\circ \text{ C. and } 760 \text{ mm.}$

$$0^\circ \text{ C. } = 273^\circ \text{ A; } 15^\circ \text{ C. is } 273 + 15 = 288^\circ \text{ A.}$$

Since the **temperature is to be lowered**, the **volume must be less**—viz.,

$$27 \times \frac{273}{288} \text{ litres, or } 25.6 \text{ litres at } 0^\circ \text{ C. and } 740 \text{ mm.}$$

Since the **pressure is to be raised**, the **volume will be less**—viz.,

$$25.6 \times \frac{740}{760} = 24.92 \text{ litres at } 0^\circ \text{ C. and } 760 \text{ mm.}$$

(This is called *reducing a gaseous volume to standard conditions.*)

Now, our equation $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$,

i.e., by weight: $65 + 98 = 161 + 2$,
and volume at S.T.P.: 22.4 litres

states that 22.4 litres of hydrogen at N.T.P. are liberated by 65 gms. of zinc;

\therefore 1 litre of hydrogen at N.T.P. is liberated by

$\frac{65}{22.4}$ gms. of zinc,

and 24.92 litres of hydrogen at N.T.P. are liberated by

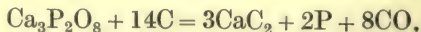
$\frac{65}{22.4} \times 24.92$, or **72.31 gms. of zinc.**

CHAPTER XXVIII

ALLOTROPY—PHOSPHORUS—OZONE

ALLOTROPY.—This phenomenon, which we investigated in the case of carbon (Expt. 165) and sulphur (Expt. 211), is in reality a fairly common one. Many of the elements are capable, under special conditions, of assuming such totally different habits and properties that they appear to be entirely different substances. Many of these changes are induced by heat (see the allotropy of sulphur); all are physical changes, involving no change in weight. It is *believed* that in all cases *allotropy is due to a difference in the number of atoms of the element that are contained in a molecule.*

PHOSPHORUS.—The element phosphorus has already been frequently met with in the preceding sections. It occurs as phosphates in many of the older rocks, and finds its way by disintegration into the soil, thence to plants and animals. The chief part of the earthy constituents of bones, to which the latter owe their rigidity, is calcium phosphate. The chief sources of phosphorus are bones (bone ash), apatite (mineral calcium phosphate), redonda phosphate (mineral aluminium phosphate), and coprolites. Calcium phosphate is decomposed by carbon in the electric furnace thus:



which is now the main process of manufacture.

The main use of phosphorus is in the manufacture

of matches. It is also the source of the oxides and halogen compounds of phosphorus—important chemical reagents. The *common* matches are tipped with a paste containing “white phosphorus” (see below), and strike on any rough surface. The use of these is discouraged, since the fumes of the phosphorus are very injurious to the “dippers.” *Safety* matches contain no phosphorus; the heads of these are composed chiefly of antimony sulphide, potassium chlorate, and glue, whilst the paste on the boxes contains “red phosphorus” (see below), manganese dioxide, or antimony sulphide and glue.

Caution.—Owing to the inflammability and poisonous character of phosphorus, experiments with it should be made by the instructor, or only under his direct supervision.

Ordinary phosphorus, when pure, is almost colourless and translucent, but as usually met with it is covered by an opaque crust; it is a **yellow, wax-like solid**, readily broken, and shows a **crystalline** structure.—After long exposure to light it becomes almost black, and very brittle.

When this substance is heated nearly to its boiling-point for a long period, it is slowly converted into a **reddish amorphous powder**, known as **red phosphorus**. This change can also be brought about thus:

EXPT.—Place a stick of phosphorus in a strong flask fitted with a cork and short tube, as shown (Fig. 49), and gradually heat it until it almost boils. Remove the flame and drop in a few crystals of iodine. Red phosphorus is produced.

EXPT.—**To compare Ordinary and Red Phosphorus.**—Besides the differences in appearance noted above, the following experiments may be performed:

1. Expose in separate dishes small quantities of these two substances, and observe that—

(1) *Ordinary phosphorus fumes*, and finally **burns** brilliantly in the air. Accordingly it is kept under water.

(2) *Red phosphorus* does **not fume nor fire**.

2. Place some red and yellow phosphorus in separate tubes, and cover with water. Put these two tubes in a water-bath and warm gently. Observe that the *yellow phosphorus melts at 44° C.*, but the *red phosphorus* does **not melt**.

3. Weigh a small graduated cylinder half full of water, and note the level of the water. Introduce a dry stick of phosphorus; reweigh, again note the level of water and calculate its density. Repeat this experiment, using red phosphorus, and observe that—

(1) *Yellow phosphorus* has a **density of 1.84 gms.** per c.cm.

(2) *Red phosphorus* has a **density of 2.15 gms.** per c.cm.

4. Place a *small* piece of yellow phosphorus in carbon disulphide and pour the clear liquid on to a filter-paper. Expose this to the *outer* air. Repeat, using red phosphorus, and observe—

(1) *Yellow phosphorus* is **soluble** in carbon bisulphide.

(2) *Red phosphorus* is **insoluble** in carbon bisulphide.

5. Warm a *small* piece of *ordinary phosphorus* in a solution of caustic soda, and observe the evolution of a gas, which ignites and has an offensive odour—it is **hydrogen phosphide (phosphine)**—and the *yellow phosphorus* **dissolves**. Repeat this experiment with red

phosphorus, and note that no phosphine is produced; the *red phosphorus* is insoluble.

The red phosphorus prepared in the first method above is purified from the yellow phosphorus by grinding it to powder under caustic soda solution.

3. Draw a slow current of air over phosphorus in a hard-glass tube (Fig. 44) in a dark room, and observe—

(1) *Yellow phosphorus* is **phosphorescent**—luminous in the dark.

(2) *Red phosphorus* is **not phosphorescent**.

Yellow phosphorus may be boiled (280° C.) in an inert gas, carbon dioxide, or nitrogen, etc., and sublimed *in vacuo* (40° C.) in brilliant crystals. *Yellow phosphorus* is very **poisonous**, whereas *red phosphorus* is said to be **inert**.

As above, yellow phosphorus can be converted into red phosphorus by heat, and we will next show that red phosphorus can be reconverted into the ordinary form.

EXPT.—To convert Red Phosphorus into White Phosphorus.—Heat about 0.25 gm. of red phosphorus in a hard-glass tube, and observe that it becomes luminous, while drops of a yellow liquid condense on the cool sides of the tube. These drops solidify on cooling, and on examination are found to be ordinary white phosphorus.

It is obvious from the above that phosphorus can exist in **different physical forms**, *mutually convertible*. When phosphorus is dissolved in molten lead, it separates on cooling in the form of dark crystals, with a density of 2.34.

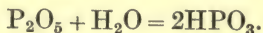
EXPT.—To show the Identity of the Allotropes of Phosphorus.—1. Repeat the Expt. on p. 116, and test the product (1) with water, (2) with litmus. In both all

the substance is *completely consumed*, and *phosphorous pentoxide alone* formed.

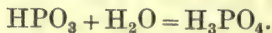
2. Repeat Expt. 165, using weighed quantities of each form. In both cases 1 gm. of either form produces 2.29 gms. of phosphorous pentoxide; so that the **two different physical substances consist entirely of the same chemical element.**

PHOSPHOROUS PENTOXIDE (PHOSPHORIC ANHYDRIDE). Molecular Formula, P_4O_{10} .—This substance was prepared on p. 116, and examined. The oxide produced there may be purified by passing its vapour, together with oxygen, over platinized asbestos. Since its vapour density is 142, its correct formula is P_4O_{10} , although the simpler formula is often employed. We found it to be a **white volatile solid, extremely hygroscopic and deliquescent, dissolving in water to form phosphoric acid.** It is our most valued **drying agent**, and is the **dehydrating agent** employed in the preparation of acid anhydrides (see Expt. 202).

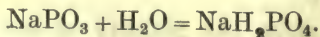
PHOSPHORIC ACIDS.—When phosphorous pentoxide is dissolved gradually in cold water or exposed to the air, it forms a glassy solid, **glacial phosphoric acid**, or **metaphosphoric acid**—



This is a **monobasic acid**, which, when boiled, absorbs water, forming *orthophosphoric acid*, thus:

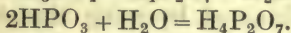


Its salts behave in a similar manner, forming *orthophosphates*—



Orthophosphoric Acid (H_3PO_4) is a **tribasic acid**, and its salts were prepared on p. 186.

A third phosphoric acid is known—viz., **pyrophosphoric acid** ($\text{H}_4\text{P}_2\text{O}_7$), which is obtained, as its name infers, by heating phosphoric acid (ortho), or by boiling metaphosphoric acid solution, thus:



This acid is **tetrabasic**.

ALLOTROPIC OXYGEN (OZONE).

Vapour density, 24; molecular weight, 48.

EXPT. 293.—To ozonize Air by Means of Phosphorus.
—Place a stick of phosphorus in a gas jar, with one end partly immersed in water. Observe the peculiar odour of the contents of the jar. Hold in this jar a strip of paper soaked in a starch solution containing potassium iodide. Observe the production of a blue colour, although a similar paper exposed to the air remains unaffected. This active, odorous gas is **ozone**, and was first noticed in 1785 by Van Marum in the vicinity of electric machines.

Small quantities of this gas are produced in many processes of slow oxidation, and in many reactions in which oxygen is liberated. The salubrity of country air is believed to be due to this gas. It is best prepared for experimental purposes as below:

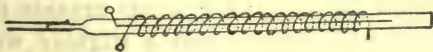


FIG. 96.

EXPT. 294.—To prepare “Ozone.”—The tube shown (Fig. 96) has a platinum wire running down the central axis, and a second wire (platinum preferably, although copper may be used) is wound spirally around the out-

side. These wires are connected to a small Ruhmkorff's sparking coil, and *dry* oxygen sent *slowly* through the tube. The smell of ozone is at once observed, and this

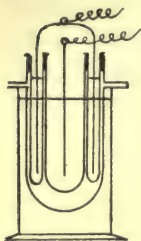


FIG. 97.

ozonized oxygen may be used for the experiments below. A better yield is obtained by using the simple apparatus shown in Fig. 97. A U-tube, provided with side-tubes, has cemented into each limb with paraffin wax a test-tube of slightly smaller bore. These test-tubes are filled with ice-cold brine or dilute sulphuric acid, and the apparatus is placed in a beaker of the same. The

wires from the coil dip into the test-tubes and outer vessel respectively, and **pure dry oxygen** is *slowly* sent through the apparatus.

Pure ozone is obtained from ozonized oxygen by passing the latter through a tube immersed in boiling liquid oxygen. The gas condenses to an *intense blue, explosive, liquid* (b.-pt., -106°C.), whilst the oxygen escapes.

EXPT. 295.—To examine the Properties of Ozone.—As we have learned above, it is a **colourless gas**, with a **peculiar odour**.

Ozone is a most powerful oxidizer.

1. Attach a small piece of india-rubber tubing to the exit tube, and notice how quickly it falls into decay. **Ozone attacks and destroys organic matter.**

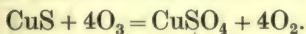
For future experiments slip a slightly wider tube over the exit side-tube and cement it in position with paraffin wax.

2. Place some *clean* quicksilver—*i.e.*, mercury that runs without a “tail”—into a *clean* flask, and pass into the latter ozonized oxygen. Roll the mercury gently over the flask; it adheres, forming a “mirror,”

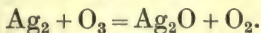
and on being shaken vigorously, a dirty grey powder is obtained. This is due to the **oxidation** of the mercury, and is a *delicate test* for ozone. **Ozone rapidly acts on most metals.**

3. Pass some ozone through a little indigo solution; it is rapidly bleached. **Ozone bleaches vegetable colours.**

4. Expose a filter-paper soaked with a solution of copper sulphate or lead nitrate *first* to hydrogen sulphide (black copper sulphide or lead sulphide is produced), and then to the action of ozone. The black sulphide is gradually *oxidized* by the ozone.



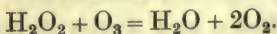
5. Pass ozonized oxygen through a hard-glass tube containing reduced metallic silver, and test the issuing gas; it is free from ozone, and silver oxide remains.



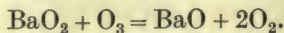
6. Pass ozonized oxygen through a tube containing silver oxide, and test the escaping gas. It is free from ozone, for the action represented by the equation $\text{Ag}_2\text{O} + \text{O}_3 = \text{Ag}_2 + 2\text{O}_2$ occurs.

Processes 5 and 6 go on simultaneously—*i.e.*, **alternate oxidation and reduction**. Many other vigorous oxidizers behave in the same way—*e.g.*, the oxides of manganese and copper, etc., hydrogen peroxide, barium peroxide, potassium permanganate, etc., both ozone and the oxidizing agent being **reduced**. In reality, the *nascent oxygen atom* (*q.v.*) is concerned in oxidizing a second oxygen atom, and gives us *good evidence in favour of the molecule* O_2 .

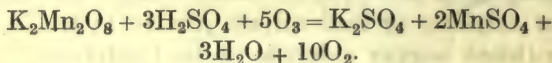
7. Bubble ozonized oxygen through hydrogen peroxide, and note that the odour disappears. The action is expressed by the equation—



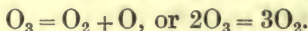
8. Pass ozonized oxygen over barium peroxide in a hard-glass tube, and test the escaping gas. It no longer contains ozone.



9. Bubble ozone through acidified potassium permanganate, and notice that the colour is discharged.

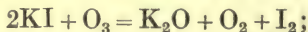


10. Pass ozone through a piece of glass tubing, heated in the Bunsen flame, and test the issuing gas. It is no longer ozone, but oxygen, for at 250° to 300° C. ozone is decomposed; thus:

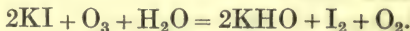


It is due to the ease with which ozone liberates **nascent** oxygen that it owes its powerful oxidizing action.

11. Bubble ozonized oxygen through a solution of potassium iodide; the brown colour of free iodine is at once observed. The nascent oxygen replaces the iodine thus:



or in aqueous solution:



If starch solution is also present, the characteristic blue colour is at once obtained. This reaction is used to *determine the amount* of ozone, an acidified (Why?) solution of potassium iodide being employed.

THE COMPOSITION OF OZONE.—1. Drs. Andrews and Tait showed that when oxygen was ozonized a contraction occurred. When this ozonized oxygen was heated, the original volume of oxygen was regained, and the process could be repeated indefinitely.

2. When Andrews and Tait's experiment was re-

peated, but with a small thin glass bulb of potassium iodide solution in the apparatus, on passing the discharge, a contraction was observed. On shaking the apparatus, the bulb was broken, the ozone was destroyed, but no further contraction occurred. Thus the ozone liberated its own volume of oxygen.

From this it was easy to *deduce* that the molecule of ozone was O_3 . This was confirmed by Soret, who absorbed the ozone without decomposition by means of turpentine, etc.

EXPT. 296.—To show the Absorption of Ozone by Means of Turpentine.—Fill the apparatus shown with ozonized oxygen by displacement, and quickly fix in the cork through which it passes a manometer-tube (*m*) and a “filler” (*b*) filled with turpentine. Place a beaker of water under *m*, and gently squeeze *b*. Turpentine enters and absorbs the ozone, as is indicated by the rise of the liquid in *m*.



FIG. 98.

3. Soret enclosed a bulb of turpentine in the apparatus used by Andrews and Tait, and found a first contraction—**a**—occurred when ozonization took place, and a second contraction—**2a**—took place after the bulb was broken, due to the absorption of the ozone. Thus $3a$ volumes of oxygen gave $2a$ volumes of ozone, or $3O_2 = 2O_3$.

This has been confirmed by diffusion experiments in comparison with chlorine and carbon dioxide, and finally Ladenburg has determined the vapour density of pure ozone to be 24. Its molecular weight is therefore 48, or 3×16 , and its formula O_3 .

REVISION QUESTIONS.

CHAPTER XXV.

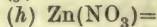
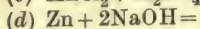
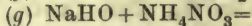
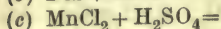
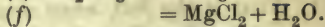
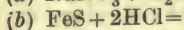
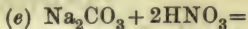
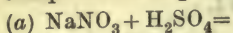
1. Two oxides of an element contain respectively 22.54 and 36.80 per cent. of oxygen. Show that the combinations of the element with oxygen are in multiple proportions. [C.L.J., 1903.]
2. State the law of multiple proportions, and suggest some simple experiment by which the law might be illustrated in the laboratory. A certain salt forms two crystalline compounds, which have the following percentage composition:
Anhydrous salt in (a), 63.86; in (b) 89.83.
Water in (a) 36.14; in (b) 10.17.
Show how these numbers illustrate the law. [L.M., 1912.]
3. Show that the following results of analyses agree with the law of multiple proportion:
(a) Iron, 63.63 per cent.; sulphur, 36.37 per cent.
(b) Iron, 46.67 per cent.; sulphur, 53.33 per cent.
4. The two chlorides of mercury on analysis yielded the following results:
(a) Mercury, 73.80 per cent.; chlorine, 26.20 per cent.
(b) Mercury, 84.93 per cent.; chlorine, 15.07 per cent.
Show that these results are in harmony with the law of multiple proportions.
5. State and illustrate the law of combination of gases by volume. What weight of oxygen would be required for the complete combustion of 500 c.cms. of carbon monoxide at 0° C. and 760 mm. pressure? [1 litre of hydrogen at 0° C. and 760 mm. pressure weighs 0.09 gm.] [C.L.J., 1907.]
6. What is Avogadro's hypothesis, and what important consequences follow from its use? Give the steps in the proof that a molecule of hydrogen consists of two atoms, clearly pointing out the experimental and theoretical steps.
7. What do you understand by a "law," a "theory," and a "hypothesis"? Clearly point out the differences, if any, between them.

CHAPTER XXVI.

1. What is the *full, exact* meaning of the following: O , H_2 , $3Zn$, O_3 , HgS , $CuSO_4$, $Pb.2NO_3$, $MgCl_2.6H_2O$, $2FeS_2$, $Al(OH)_3$, $K_2SO_4.CuSO_4.6H_2O$?
2. Give the molecular formulæ and the densities (relative to hydrogen) of the following substances in the state of gas: Chlorine, nitric oxide, ozone, hydrogen sulphide.
3. Write down any one chemical equation you know, and state *all* that can be inferred from it. [C.P., 2nd, 1906]
4. Calculate the percentage compositions of the following:
 - (a) Nitric acid, HNO_3 ; (b) magnesium carbonate: $MgCO_3$; (c) hydrochloric acid, HCl ; (d) crystallized ferrous sulphate, $FeSO_4.7H_2O$; (e) sodium hydrate, $NaHO$; (f) potassium hydrogen sulphate, $KHSO_4$.
5. Find the percentage composition of normal sodium sulphate. [$O=16$, $Na=23$, $S=32$.] [C.L.J., 1903.]
6. Calculate the simplest formula from the following percentage compositions:
 - (a) Carbon, 42.86 per cent.; oxygen, 57.14 per cent.
 - (b) Sodium, 43.40 per cent.; carbon, 11.32 per cent.; oxygen, 45.28 per cent.
 - (c) Zinc, 40.37 per cent.; oxygen, 39.75 per cent.; sulphur, 19.88 per cent.
 - (d) Calcium, 40 per cent.; oxygen, 48 per cent.; carbon, 12 per cent.
 - (e) Nitrogen, 35 per cent.; hydrogen, 5 per cent.; oxygen, 60 per cent.
 - (f) Zinc, 22.65 per cent.; sulphur, 11.15 per cent.; oxygen, 61.32 per cent.; hydrogen, 4.88 per cent.
 - (g) Copper, 47.01 per cent.; chlorine, 52.99 per cent.;
 - (h) 46.67 per cent. of iron; 53.33 per cent. of sulphur.
7. Calculate the formula of a substance from the percentage composition: Calcium, 47.90; phosphorus, 18.56; oxygen, 33.54. [$O=16$, $P=31$, $Ca=40$.] [C.L.J., 1908.]
8. Explain the meaning of the following equations:
 - (a) $NH_3 + HCl = NH_4Cl$;
 - (b) $Ca(HO)_2 + CO_2 = CaCO_3 + H_2O$.

[$N=14$, $Cl=35.5$, $Ca=40$, $C=12$, $O=16$.] [C.L.J., 1909.]

9. Complete and explain the following equations:



10. Write down the equations for the changes that occur, and determine from them—

(a) The weight of oxygen obtained by heating 70 gms. potassium chlorate.

(b) The weight of red lead required to yield 20 litres of oxygen at S.T.P.

(c) The weight of magnesium required to set free 5 lbs. of hydrogen.

(d) The weight of iron and steam required to produce 50 litres of hydrogen at S.T.P.

(e) The weight of limestone required to make 10 tons of lime.

(f) The weight of dry copper nitrate required to make 5 lbs. of copper oxide.

(g) The weight of sulphuric acid to yield 50 litres of sulphur dioxide at N.T.P.

(h) The weight and volume of hydrogen chloride to produce 5 litres of chlorine at S.T.P.

(i) The weight and volume of ammonia gas at S.T.P. required to produce 15 gms. of ammonium chloride.

(j) The volume and weight of oxygen and hydrogen (at S.T.P.) obtainable by electrolyzing 60 gms. of water.

(k) The weight and volume of sulphur dioxide (at S.T.P.) obtained by roasting 200 gms. of iron pyrites.

(l) The weight of iron sulphide and the volume of hydrogen sulphide (at S.T.P.) obtainable from 20 gms. of sulphur.

11. What volume of hydrogen and oxygen may be obtained at a temperature of 15°C . when 3 gms. of water are decomposed? [11.77 litres of hydrogen at 15°C . weigh 1 gm.] [C.L.J., 1909.]

12. Find the weight of (a) potassium chloride to convert 150 gms. of lead nitrate into lead chloride, and of the latter formed; (b) caustic soda to convert 70 gms. of copper sulphate crystals into copper oxide, and of the oxide produced; (c) sulphuric acid to produce 50 gms. of sodium sulphate (anhyd.) from washing-soda.

CHAPTER XXVII.

1. State Boyle's law, and give an experimental proof.
2. Calculate the volume of—
 - (a) 10 litres of oxygen at 15° C. and 740 mm., when measured at 15° C. and 600 mm.
 - (b) 5 litres of hydrogen at 25° C. and 750 mm., when measured at 25° C. and 850 mm.
 - (c) 200 c.cms. of carbon dioxide at 0° C. and 760 mm., when measured at 0° C. and 550 mm.
 - (d) 50 c.cms. of ammonia at 10° C. and 740 mm., when measured at 10° C. and 760 mm.
3. State the law of Charles. Describe an experiment to illustrate it.
4. How would you demonstrate that a gas expands when it is heated, and contracts when it is cooled? A mass of carbon monoxide occupies a volume of 273 c.cms. at 17° C. What will its volume be at 32° C., the pressure remaining constant? [C.P., 2nd, 1906.]
5. What do you understand by "absolute temperature," and how is it related to the centigrade scale?
6. Determine the volume of—
 - (a) 25 cubic inches of nitrogen at 0° C. and 760 mm., when measured at 25° C. and 760 mm.
 - (b) 52 litres of hydrogen at 15° C. and 800 mm., when measured at 0° C. and 800 mm.
 - (c) 16 c.cms. of "laughing-gas" at 10° C. and 740 mm., when measured at 35° C. and 740 mm.
 - (d) 1 cubic foot of oxygen at 25° C. and 700 mm., when measured at 15° C. and 700 mm.
7. Calculate the volume of—
 - (a) 20 litres of air at 0° C. and 760 mm., when measured at 10° C. and 800 mm.
 - (b) 15 cubic feet of hydrogen at 15° C. and 740 mm., when measured at 0° C. and 760 mm.
 - (c) 10 c.cms. of ammonia at 10° C. and 780 mm., when measured at 30° C. and 640 mm.
 - (d) 15 litres of carbon dioxide at 10° C. and 800 mm., when measured at 0° C. and 760 mm.
8. What would be the weight of the solid product obtained if 200 c.cms. of ammonia, measured at 17° C. and

75 cms. pressure were mixed with an equal volume of hydrogen chloride, measured at the same temperature and pressure ?

9. What weight of water would have to be decomposed in order to supply the necessary hydrogen to completely combine with 1 litre of chlorine, measured at 20° C. and 750 mm. of mercury pressure ? [C.P., 2nd, 1911.]
10. What volume of oxygen would be obtained at 25° C. and 790 mm. pressure from 50 gms. of mercuric oxide ?
11. Calculate the weight of sulphuric acid required to furnish 50 litres of hydrogen at 15° C. and 750 mm. pressure ?
12. Determine the volume of carbonic acid gas obtained at 20° C. and 600 mm. pressure from 1 kilogramme of marble.
13. What volume of carbon dioxide at 15° C. and 740 mm. pressure will be required to convert 60 gms. of lime into chalk ?
14. What is the volume of 200 c.cms. of air at 16° C. and 770 mm. pressure when reduced to S.T.P. ?
15. What volume of ammonia gas at 10° C. and 730 mm. pressure will be required to neutralize 50 gms. of nitric acid, and what will be the weight of the product ?
16. What volume of nitrogen will be obtained at 16° C. and 745 mm. pressure by heating 25 gms. of ammonium nitrite ?
17. Calculate the volume of ammonia obtainable from 70 gms. of ammonium nitrate when measured at 12° C. and 770 mm.
18. What volume of sulphur dioxide will be obtained at 20° C. and 747 mm. pressure by roasting 5 kilogrammes of iron pyrites ?
19. What volume of chlorine is necessary to decompose 14 gms. of hydrogen sulphide at 15° C. and 740 mm. pressure ?
20. Find the volume of hydrogen chloride produced when 50 gms. of common salt are decomposed by sulphuric acid at 12° C. and 765 mm. pressure.
21. Determine the volume of chlorine at 20° C. and 740 mm. liberated by 100 gms. of manganese dioxide.
22. Calculate the volume and weight of sulphur dioxide required to decompose 50 gms. of hydrogen sulphide at 15° C. and 750 mm. pressure.
23. What volume of nitric oxide at 15° C. and 770 mm. is required to completely oxidize 5 gms. of copper ?

24. What volume of nitrogen peroxide (NO_2) at 10°C . and 700 mm. is required to completely oxidize 20 gms. of mercury ?
25. What weight of hydrogen chloride is required to produce 500 litres of chlorine at 12°C . and 750 mm. pressure, when heated with lead peroxide ?

CHAPTER XXVIII.

1. What do you understand by "allotropy" ? Illustrate your answer by reference to carbon and sulphur.
2. To what is allotropy due ? Give experimental evidence in support of your statement.
3. How would you prove that phosphorus exhibits allotropy ?
4. What is ozone ? How is it prepared ?
5. What evidence have we that the formula of ozone is O_3 ?
6. Describe experiments in which ozone acts as an oxidizer.
7. "Ozone is an allotropic form of oxygen." What does this mean, and how would you justify this statement ?
8. Mention the chief properties of ozone, and illustrate them by experiments.
9. How would you convert red phosphorus into yellow, and reverse the process ? Mention properties in which these differ, and any uses made of each.
10. How would you prepare from phosphorus (a) phosphoric acid, (b) sodium phosphate ? How could you regain the phosphorus from the former ?

EXAMINATION PAPERS.

OXFORD LOCAL EXAMINATIONS.

JULY, 1909. JUNIOR CHEMISTRY. (TWO HOURS.)

[Draw sketches of the apparatus you mention, and give equations where possible.]

1. What happens when a piece of metallic sodium is put into water ? Given the atomic weights, $\text{Na} = 23$, $\text{H} = 1$, and $\text{O} = 16$, it is possible to calculate the *volume* of one of the products obtainable from a given weight of sodium. Explain this.

2. Nitrous oxide supports combustion almost as well as oxygen itself. Describe three experiments by which, given a supply of each of these gases, you could distinguish the one from the other.
3. The names given to substances by chemists are usually intended to show what those substances are made of. Illustrate this by giving and explaining the chemical names of nitre, quicklime, slaked lime, washing-soda, and bone-ash.
4. Describe carefully some experiment by which it can be proved that a diamond and a piece of "blacklead" (or graphite) consist of the same substance.
5. It is stated that air contains carbon dioxide. How could you prove this, and how do you account for the presence of that gas in air? What has the fact of its being there to do with the properties of many natural waters?
6. Name four instances of chemical decomposition produced by heat, and give the equation representing each of them.

Practical—A.

1. The powder *X* is a metal. When it is heated in contact with air, it is converted into oxide. You are asked to find out what weight of *X* would combine with 16 gms. of oxygen. Use for this experiment about 2 gms. of *X*; heat it to redness for half an hour or more; you can be doing Question 2 meanwhile.
2. Shake up 1 to 2 gms. of the powder *Y* with about 100 c.cms. of cold distilled water for several minutes, and then filter the clear liquid. Find out, by experiments upon the clear filtrate (*a*) whether it contains any dissolved matter; (*b*) whether it is neutral, acid, or alkaline; and (*c*) whether any visible change occurs when a portion of it is boiled in a clean test-tube.

B.

1. The powder *Z* is the oxide of a metal. Find out, by analysis, what that metal is. Perform and describe fully an experiment to show that *Z* is an oxide.
2. The envelope marked *a* contains some pieces of a metal. You are asked to find out what happens when separate pieces of it are treated with (*a*) strong nitric acid;

- (b) strong nitric acid diluted with an equal volume of water; (c) (bench) dilute nitric acid; (d) (bench) dilute nitric acid diluted with three times its volume of water.
3. Find out what is the acid of which the substance β is a salt, stating whether the envelope is marked β_1 , β_2 , or β_3 .

JULY, 1911. JUNIOR CHEMISTRY. (TWO HOURS.)

[Draw sketches of all the apparatus you mention, and give equations where possible.]

1. Describe an experiment to prepare nitrogen from the air by the use of copper. What common compounds does nitrogen form with hydrogen and with oxygen? Show how two of them are prepared.
2. Why is it believed that (a) steam is a chemical compound of hydrogen and oxygen, (b) oxygen is an element?
3. What is Avogadro's law, and how does it help to determine the relative weights of gaseous molecules?
4. Write out the formula of a ferrous and of a ferric salt. State how it is possible to convert a ferrous into a ferric salt. What is iron pyrites, and to what use is it put?
5. How is nitric acid obtained from saltpetre? Describe the apparatus used, and state what is the appearance of the pure acid. What takes place when it is (a) boiled, (b) added to marble, (c) warmed with copper?
6. In what way is pure sulphur dioxide obtained? Describe an experiment by which the composition by volume may be determined.
7. Explain why a greater yield of hydrochloric acid may be obtained from sodium chloride than from an equal weight of potassium chloride.

JUNIOR PRACTICAL CHEMISTRY. (TWO HOURS.)

A Paper.

1. Weigh an evaporating-dish of medium size, and place in it about 2 gms. of the wire I. with which you are provided. Weigh the whole again. Cover the dish with a watch-glass; add dilute nitric acid, and warm gently until the whole is dissolved. (Add a drop or two of concentrated nitric acid if solution is slow.) Then evaporate very

carefully to dryness. Finally heat strongly until no further change takes place. Allow to cool, and weigh again. Calculate from the results the weight of the metal which combines with 8 parts by weight of oxygen. Describe fully all the changes which take place during the experiment.

2. The substance marked II. is either a chemical compound or a mixture of two chemical compounds. Determine which it is by the following experiments: (a) Dissolve about half the substance in warm water; filter, and dry any residue, and evaporate the filtrate in a basin. (b) Describe the residues left behind upon the filter-paper and on evaporation. To each residue add a little dilute hydrochloric acid, and then warm a fresh portion of each residue with caustic soda solution. State what occurs in all your experiments. What elements are contained in the substance II. as determined by these experiments?

B Paper.

1. Determine the nature of the substances marked III. and IV. Dissolve a mixture of about equal weights of III. and IV. in dilute sulphuric acid, and gradually heat to boiling in a flask fitted with a gas delivery-tube. Boil for a few minutes and identify (a) the gas evolved during the heating, and (b) the nature of the metallic radicals in the solution in the flask after the experiment.
2. Determine whether the substance V. is an oxidizing or reducing substance; describe exactly the experiments which you perform.

CAMBRIDGE LOCAL EXAMINATIONS.

1912.—JUNIOR CHEMISTRY. (TWO HOURS.)

[Only *six* questions may be attempted.]

1. State the law of definite proportions, and describe experiments by which you could verify it in some particular case.
2. Describe in detail how the composition of water *by weight* may be determined experimentally.

3. Explain what is meant by *oxidation*. Mention two important oxidizing agents, and give examples to illustrate their action.
4. Describe the preparation of carbon dioxide. A piece of charcoal is burnt in oxygen; by what tests would you show that carbon dioxide is formed, and how could you recover some of the carbon ?
5. Describe a method for the preparation of nitrogen from sources other than air. What proofs are there that the nitrogen in the atmosphere is mixed and not combined with the other chief component of the air ?
6. Calculate the volume of oxygen obtained by strongly heating 25 gms. of potassium chlorate, having given that 1 litre of hydrogen weighs 0.084 gm. at the temperature and pressure at which the oxygen is measured. [K = 39, Cl = 35.5, O = 16.]
7. What changes occur when the following substances are heated in air: (a) Lead, (b) lead peroxide, (c) calcium carbonate, (d) ammonium nitrate ?
8. Describe carefully an experiment to illustrate Boyle's law. Why is the pressure of air greater at sea-level than at the top of a mountain ?

PRACTICAL CHEMISTRY. (TWO HOURS.)

[Only *two* questions to be answered.]*A. Paper.*

1. Estimate the loss in weight of the substance *M1* after it has been strongly heated. You should repeat the process until the weight is constant. Calculate your results for 100 gms. of *M1*.
2. Heat the substance *N1* with concentrated sulphuric acid, and describe what you observe. Identify the gas which is given off, and confirm your results by as many tests as you can apply.
3. Identify the substance *O1* by as many tests as you can.

B. Paper.

1. Estimate the volume of the gas given off when 0.05 gm. of the metal *P1* is dissolved in dilute hydrochloric acid. State your result corrected to 0° C. and 760 mm. pressure.

2. Heat the substance *D1* with concentrated hydrochloric acid, and describe what you observe. Identify the gas which is given off, and confirm your observation by as many tests as you can apply.
3. Identify the substance *E1* by as many tests as you can.

THE COLLEGE OF PRECEPTORS.

SECOND CLASS. CHEMISTRY (A). ($1\frac{1}{2}$ HOURS.)

[Not more than *five* questions to be attempted.]

Christmas, 1911.

1. Distinguish fully between a *physical* and a *chemical* change, illustrating your answer by reference to the behaviour of red lead when heated.
2. Sketch and describe an apparatus for illustrating *diffusion* of gases. What experiments would you carry out with the apparatus to demonstrate this property?
3. Write an essay, not exceeding two pages, on combustion and describe at least one experiment in which the air is not to be regarded as the "supporter of combustion."
4. How would you proceed to prepare a specimen of at least 10 c.cms. of concentrated nitric acid? Describe the obvious properties of the specimen and experiments for characterizing the acid.
5. Give examples of substances which are known in two or more physically distinct forms, and proofs that such forms are chemically identical.
6. You are required to estimate the strength of a sample of washing-soda. Describe in detail how you would proceed.
7. How would you prepare a sample of pure nitrous oxide, and how would you test it for the presence of nitric oxide? Compare the properties of nitrous oxide with those of oxygen, and show how these gases may be readily distinguished.
8. Describe methods of obtaining from concentrated nitric acid, specimens of (a) hydrogen, (b) nitrogen, (c) oxygen.

Midsummer, 1912.—Section A.

1. Fully describe the method you would adopt to obtain pure hydrogen from a mixture of equal volumes of this gas and carbon dioxide.
2. Describe the preparation of a strong aqueous solution of ammonia. How would you prove that ammonia contains nitrogen and hydrogen?
3. Write a short essay on combustion.
4. What simple experiments would you make to prove the presence of quicklime and chalk in a mixture of these compounds?
5. How would you prepare ozonized oxygen? How would you prove the presence of ozone in it? Give a short account of the chief properties of ozone.
6. Starting with sodium carbonate, describe the preparation of (a) caustic soda, (b) sodium sulphate, (c) sodium bisulphate, and write equations for the reactions involved.
7. Copper is (a) heated in air, (b) heated with strong sulphuric acid, (c) treated with nitric acid. Describe all you would expect to observe in each case, and state the nature of the changes which occur.
8. 50 gms. of pure calcium carbonate are heated until the residue has a constant weight. If the evolved gas were measured at 0° C. and 760 mm. pressure, what volume would it occupy? What would be the weight of the solid residue? [Atomic weights: Ca = 40, O = 16, C = 12.]

Christmas, 1912.

1. Define the terms "atom," "molecule," "atomic weight," and "molecular weight." Explain the numerical relationship between the molecular weight and the density of oxygen referred to hydrogen.
2. Describe in detail the preparation of one metallic salt of nitric acid other than that of an alkali metal.
3. How would you prove that copper oxide is composed wholly of copper and oxygen?
4. Sketch and describe an apparatus suitable for the preparation of a concentrated aqueous solution of hydrogen

chloride from common salt. How can the anhydrous gas be obtained from its solution ?

5. Describe in detail how the composition of water by weight can be experimentally determined. Sketch the apparatus used ?
6. In what ways can sulphur dioxide be converted into sulphuric acid ? Sketch and describe an apparatus illustrating one of the methods.
7. How would you prepare chlorine, and with what experiments would you demonstrate its most important properties ?
8. Describe experiments illustrating the use of each of the following as reducing agents, and write equations to express the changes which take place in each case: Carbon, hydrogen, sulphur dioxide, and carbon monoxide.

APPENDIX

NOTES ON PRACTICAL WORK

I. *The Observation of the Effects of Heat, Water, or Dilute Acids, etc., upon Materials to be examined.*

1. Follow the instructions given in the question *carefully and exactly.*

2. When examining the effect of heat on a substance, use a small ignition-tube held *horizontally* in the flame, and note—

(a) Any changes in colour, melting, etc., that occur.

(b) (1) Any *liquid* that collects on the upper portions of the tube. Test it with litmus-papers. (Does it solidify when cold?) Observe its appearance. (Is it sulphur, mercury, etc.?)

(2) Any *solid sublimate* formed on the cool sides of the tube. Observe its appearance. (Is it quicksilver, sulphur, ammonium salt, etc.?)

(c) Any *gas or vapour* that escapes. Observe its odour, colour, action on litmus-papers. Test it with a lighted taper, a glowing wood ember, a spot of lime-water held on a glass rod, lead acetate paper, paper soaked in acidified potassium bichromate. Many gases—*e.g.*, hydrogen, oxygen, carbon dioxide, sulphuretted hydrogen, sulphur dioxide, nitrogen peroxide, etc.—may be readily identified in this way. Collect a little of the gas in test-tubes over water (Fig. 56). Test the gas, and *test the water in the dish.*

- (d) Any *residue* that remains. Compare it with the original substance, and with substances you know well. Test it with litmus-papers, water, and dilute acids.
3. Similarly, when observing the effect of water and dilute acids upon a substance, note—
- (a) Any change in colour, etc., that takes place.
- (b) Any gases that escape, and test them as above. Fit your tube with a cork and delivery-tube. Collect the gas over water, and test as before. When the action is over, filter the residual liquid and evaporate it. Examine the residue, if any, and compare it with the original substance.
- (c) Any residue that remains. (Is it the same, or different from the original substance ?)

Unless your instructions forbid, when using water and acids, use them cold and warm, and the latter weak and strong, and observe any differences in action that may take place. Most of the substances studied will be readily detected, and from the information gained you will be in a position to draw an intelligent conclusion to your work, which will greatly add to the interest of it and its value.

II. *When required to Identify a Simple Substance* the above tests will be of great assistance. It is well, also, to have a good knowledge of the commoner materials. Examine those on the laboratory shelves *frequently* and those you prepare yourself. Write brief descriptions of them from memory, and compare them with the originals until these become more than mere acquaintances to you.

III. You may be required to apply *confirmatory* tests to verify your conclusions as to the presence of a particular acid or metal in the substance supplied, and in all cases it is *wiser* to do so.

A. TO CONFIRM THE PRESENCE OF AN ACID.

1. To the original substance add a little dilute hydrochloric acid (or sulphuric acid), and warm, if necessary.

Carbonates (see Expt. 125) are decomposed with effervescence, and carbon dioxide is liberated. This may be identified by its action on lime-water.

Sulphites (see Expt. 191) set free sulphur dioxide, which may be recognized by its pungent odour, its action on litmus, and on acidified potassium bichromate.

Nitrites (see Expt. 220) liberate brown fumes. These may be tested by adding cold ferrous sulphate solution, and obtaining the brown layer.

2. Add a little *strong* sulphuric acid to the substance, and *gently* warm.

Nitrates (Expt. 216) yield nitric acid. Its presence may be readily detected by (1) adding a small piece of copper, and obtaining the brown fumes; or (2) adding to the cold mixture a cold solution of ferrous sulphate, and obtaining the brown layer.

Chlorides (Expt. 243) generate hydrochloric acid gas, easily recognized by its steamy acid fumes. These may be confirmed (1) by holding a glass rod dipped in ammonia solution near the tube; (2) by adding a little manganese dioxide, and warming; the greenish-yellow, suffocating gas, chlorine, which bleaches litmus, is obtained.

Oxalates (Expt. 155) are decomposed, a mixture of carbon monoxide and carbon dioxide being set free. The former may be identified by applying a light, the latter by its action on lime-water.

Oxygen liberated here denotes the presence of an oxidizing agent, and may be confirmed by warming the substance with strong hydrochloric acid.

3. **Sulphates** may be recognized as in Expt. 187, by adding to the solution of the substance in water a few drops of hydrochloric acid, and then barium chloride. An insoluble white precipitate of barium sulphate indicates the presence of a sulphate.

B. TO CONFIRM THE PRESENCE OF A METAL.

Since certain metals always form coloured salts, etc., a knowledge of these will often prove useful—*e.g.*, copper salts are green or blue, their solutions being blue; manganese forms brown or pink salts; ferrous salts are pale green; ferric salts yellow or white; and so on.

Since most metals form insoluble hydroxides (or carbonates) (see Expt. 180), we may use them in our identification.

I. To a solution of the substance in water or dilute acid add sodium hydroxide.

(A) **If no precipitate** is produced, Pb, Cu, Fe, Ca, Mn, Zn, Mg are absent. Test the original substance for—

1. **Ammonium**.—Warm the substance with caustic soda solution. An evolution of ammonia indicates the presence of ammonium.
2. **Potassium**.—Heat a small fragment of the substance on a platinum wire moistened with hydrochloric acid. A *lilac*, coloured flame indicates potassium.
3. **Sodium**.—Repeat 2; a *vivid golden yellow* flame is obtained if sodium is present.

(B) **If a precipitate** is produced, examine it thus:

1. **Copper** is indicated by a *blue* precipitate, which becomes *black* on boiling. This may be confirmed by passing sulphuretted hydrogen through the acidified solution. A black precipitate is obtained if copper is present.

2. **Iron (ferric)** is shown by a *reddish-brown* precipitate. This may be confirmed by (a) passing hydrogen sulphide through the acidified solution, when *sulphur* is precipitated. On adding ammonia to this, a *black* precipitate of ferrous sulphide is obtained. Sodium carbonate gives a *reddish-brown* precipitate.
3. **Iron (ferrous)** solutions give a *dirty greenish-white* precipitate, which turns *brown* on exposure. This may be confirmed by warming the substance with nitric acid, and then adding sodium hydrate. A *reddish-brown* precipitate is obtained. When hydrogen sulphide is passed through an acidified solution of a ferrous salt, nothing occurs, but on adding ammonia, a *black* precipitate is obtained. Sodium carbonate gives a *white* precipitate, turning *brown*.
4. **Manganese** salts give a *white* precipitate, turning *brown*. This may be verified by passing hydrogen sulphide through a solution containing ammonia. A *flesh-coloured* precipitate is formed. Sodium carbonate gives a *white* precipitate, turning *brown* on exposure.
5. **Lead** is indicated by a *white* precipitate, soluble in excess of soda. It may be confirmed by heating a little of the precipitate in a tube; *yellow* lead oxide remains. When hydrogen sulphide is passed through an acid solution, *black* lead sulphide is precipitated. On adding hydrochloric acid to a solution of lead, *white* lead chloride is precipitated, which dissolves on heating, and reappears on cooling.
6. **Zinc** also gives a *white* precipitate, *soluble* in excess of soda; but when the precipitate is heated, *white* zinc oxide remains, which is yellow while hot. This may be confirmed by passing hydrogen sulphide through a solution containing ammonia, when *white* zinc sulphide is obtained.

7. **Calcium, Magnesium.**—Solutions of these metals give a *white* precipitate, which is *insoluble* in excess of soda. The presence of calcium may be confirmed (1) by its flame test; a *brick-red* flame denotes calcium. If calcium is not present, magnesium may be confirmed by adding to its solution containing ammonia a little sodium hydrogen phosphate. On vigorously shaking, a *white crystalline* precipitate is produced if magnesium is present.

II. Having identified the metal and the acid present in the substance supplied, your results must be examined. For instance, if lead and sulphate have been found, the substance must be insoluble. Should the original substance be soluble, you must repeat your observations and check them to verify your result.

ANSWERS AND KEY.

- A. (1) MgCO_3 ; (2) white lead; (3) MnCO_3 ; (4) ZnCO_3 .
 B. (1) PbO_2 ; (2) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$; (3) MgCO_3 ; (4) KHC_2O_4 .
 C. (1) CuC_2O_4 ; (2) NaCO_2H ; (3) $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$; (4) $\text{Ca}(\text{NO}_3)_2$.
 D. (1) FeC_2O_4 ; (2) NaHCO_3 ; (3) ZnCO_3 ; (4) KHC_2O_4 .
 E. (1) $\text{Na}_2\text{CO}_3 + \text{C}$; (2) $\text{NH}_4\text{Cl} + \text{CaCO}_3$; (3) $\text{CuO} + \text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.
 (4) $\text{K}_2\text{CO}_3 + \text{S}$. F. (1) HgO (yellow); (2) Fe_2O_3 ; (3) Cu_2O ;
 (4) KClO_3 ; (5) CaCO_3 ; (6) PbO_2 ; (7) CuCO_3 ; (8) $\text{H}_2\text{C}_2\text{O}_4$.
 G. (1) CaCO_3 ; (2) PbCO_3 ; (3) BaCO_3 ; (4) Cu (pulv.);
 (5) MnO_2 . H. (1) KHC_2O_4 ; (2) NaHCO_3 . I. (1) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$;
 (2) $\text{Hg}_2(\text{NO}_3)_2$; (3) MnCl_2 ; (4) MgSO_4 . J. (1) Iceland spar;
 (2) dolomite. K. (1) Fe ; (2) Cu ; (3) Mg ; (4) Zn . L. (1) CuSO_4 ;
 (2) $\text{Hg}(\overline{\text{NO}_3})_2$; (3) PbAc_2 ; (4) SnCl_2 . M. (1) HCl ; (2) NaHO ;
 (3) KHO ; (4) H_2SO_4 ; (5) NaHO ; (6) Am_2CO_3 . N. (1) $\text{Pb}(\overline{\text{NO}_3})_2$;
 (2) FeCl_3 ; (3) $\text{Hg}(\text{NO}_3)_2$; (4) BaCl_2 ; (5) NiSO_4 ; (6) CaCl_2 .
 O.P. (1) KCl ; (2) PbCO_3 ; (3) Am_2SO_4 ; (4) $\text{Ca}(\text{NO}_3)_2$; (5)
 $\text{Na}_2\text{C}_2\text{O}_4$; (6) PbS ; (7) CuCl_2 ; (8) CaCO_3 ; (9) Na_2SO_3 ; (10)
 $\text{H}_2\text{C}_2\text{O}_4$. Q. (1) $\text{Pb}(\overline{\text{NO}_3})_2$; (2) FeS_2 ; (3) ZnCO_3 ; (4) FeC_2O_4 .

R. (1) Pb_3O_4 ; (2) Cu_2O ; (3) MnO_2 . *S.* (1) Al filings; (2) Zn dust; (3) iron filings. *T.* (1) Cu_2O ; (2) PbO_2 ; (3) ZnO . *U.* (1) Mg; (2) Zn; (3) Cu (pulv.). *V.* Cu (pulv.). *W.* FeS_2 . *X.* Cu (pulv.). *Y.* Slaked lime. *Z.* Cu_2O (pulv.). *a.* Fe foil. β . (1). Na_2HPO_4 ; (2) Am_2SO_4 ; (3) KNO_3 . *I.* Fe wire. *II.* AmCl . chalk. *III.* FeSO_4 . *IV.* KNO_3 . *V.* PbO_2 .

CHAPTER XXIV., p. 295.—1. 39·22. 2. (1) 8·02; (2) 31·54. 3. 35·45. 4. 32·54. 5. 32·68. 6. 11·95. 7. 22·79. 8. 8·94. 9. 38·71. 10. 22·78. 11. 31·80. 12. 12·12. 13. 8·16. 14. 30·7. 15. 21·05. 16. 3·01. 17. 100. 18. 6·15. 19. 12·00. 20. 7·06. 21. 103·11. 22. 52·8. 23. 35·00. 24. 30·46. 25. 48·22. 26. 14·14. 27. (1) 59·44; (2) 29·75. 28. (1) 14·112; (2) N, 63·83 per cent. O, 36·17 per cent. 29. (1) 103·11; (2) 77·33. 30. 9·00. 31. 0·138. 32. 12·00. 33. 7·969. 34. 108·32. 35. 68·80. 36. 100·11. 37. (1) 35·45; (2) 109·09. 38. 23·05. 39. 108. 40. 31·5. 41. 27·92. 42. 108. 43. 102·50. 44. 31·20. 45. 36·09. 46. 62·39. 47. 58·56. 48. 48·59. 49. 39·49. 50. 36·43. 51. (1) 39·98; (2) 70·95. 52. 143. 53. 50. 54. (1) 28; (2) 22.

N. 4, p. 301, 49·09. *O.* 5, p. 302, 9.

CHAPTER XXV., p. 356.—5; 0·36 gm.

CHAPTER XXVI., p. 357.—4. (a) H, 1·587 per cent.; N, 22·222 per cent.; O, 76·195 per cent.. (b) Mg, 28·57 per cent.; C, 14·29 per cent.; O, 57·14 per cent. (c) H, 2·739 per cent.; Cl, 97·261 per cent. (d) (1) Fe, 20·14 per cent.; S, 11·51 per cent.; O, 23·02 per cent.; H_2O , 45·32 per cent.; (2) Fe, 20·14 per cent.; S, 11·51 per cent.; O, 63·31 per cent.; H, 5·04 per cent. (e) Na, 57·50 per cent.; H, 2·50 per cent.; O, 40·00 per cent. (f) K, 28·68 per cent.; H, 0·74 per cent.; S, 23·53 per cent.; O, 47·06 per cent. 5. Na, 32·39 per cent.; S, 22·53 per cent.; O, 45·07 per cent. 6. (a) CO ; (b) Na_2CO_3 ; (c) ZnSO_4 ; (d) CaCO_3 ; (e) NH_4NO_3 ; (f) $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$; (g) CuCl_2 ; (h) FeS_2 . 7. $[\text{Ca}_4\text{P}_2\text{O}_7]$. 10. (a) 27·43; (b) 1·223 kilogrammes; (c) 60 lbs.; (d) Fe, 93·74

gms.; H_2O , 40.18 gms.; (e) 17.86 tons; (f) 11.835 lbs.; (g) 437.68 lbs.; (h) 10 litres, 16.29 gms.; (i) 6.28 litres, 4.766 gms.; (j) O, 53.33 gms., 37.33 litres; H, 6.67 gms., 74.67 litres; (k) 213.33 gms., 74.67 litres; (l) 55 gms., 14 litres. 11. H, 3.92 litres; O, 1.96 litres. 12. (a) 67.52, 125.98; (b) 22.49, 22.21; (c) 34.51.

CHAPTER XXVII., p. 359.—2. (a) 12.33 litres; (b) 4.41 litres; (c) 276.36 c.cms.; (d) 48.68 c.cms. 4. 287.10 c.cms. 6. (a) 27.28 cubic inches; (b) 49.29 litres; (c) 17.41 c.cms.; (d) 0.966 cubic foot. 7. (a) 19.696 cubic inches; (b) 13.84 cubic feet; (c) 13.05 c.cms.; (d) 15.32 litres. 8. 0.4438 gm. 9. 0.7388 gm. 10. 2.82 litres. 11. 204.6 gms. 12. 302.3 litres. 13. 26 litres. 14. 191.5 c.cms. 15. 19.19 litres; 63.49 gms. 16. 9.45 litres. 17. 20.19 litres. 18. 2,038 litres. 19. 9.995 litres. 20. 19.85 litres. 21. 28.38 litres. 22. (1) 47.07 gms.; (2) 17.61 litres. 23. 1.851 litres. 24. 2.52 litres. 25. 3.081 kilogrammes.

EXAMINATION PAPERS, pp. 365, 367.—C.L.J., 1912: 6. 7.29 litres. C.P. (Midsummer), 1912: 8. 11.2 litres; 28 gms.

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